EXPERIMENTAL SCIENCE

II CHEMISTRY

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EXPERIMENTAL SCIENCE

II. CHEMISTRY

BY

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PREFACE

THE author claims to have been one of the first to put to a complete test in the laboratory the suggestions of that "inspiring tract"—the Report presented at the Newcastle-upon-Tyne meeting of the British Association by the Committee formed to investigate the methods of teaching chemistry. He wishes to record his indebtedness to that Committee, whose efforts set teachers of subjects other than Science thinking and experimenting. The path of adventure and discovery was opened to many a boy and girl who would otherwise have trodden the damp and muddy ways of test-tube analysis.

The short time available and the ever-changing classes render strict adherence to the "heuristic" method impracticable, while many recent discoveries give us direct paths which make it tedious and unnecessary to follow the "historic" method.

The course given in this book is the result of experience. It should be completed in two years by a class (aged from 12 to 16 years) working for two hours per week.

There is constant reiteration of fundamental facts and very little revision should be necessary if the examples are systematically set and exacted for home-work.

The first chapter is intended for demonstration by the teacher. Much time will be ultimately saved if it is carefully explained.

The Junior Chemistry papers of the Oxford and Cambridge Examinations (especially those in Experimental Science), and of the College of Preceptors, the Board of Education and the Army Qualifying Examinations may be successfully taken after completing this course.

The author is glad to record his gratitude to his colleague Mr W. R. Cooper for useful advice and assistance in reading through the proof-sheets. To his colleague Mr Lawrence Caldecott he would express his sincere thanks for preparing the examples and examination papers as well as for his untiring help and valuable suggestions.

S. E. BROWN.

June 1912.

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EXPERIMENTAL CHEMISTRY

CHAPTER I.

APPARATUS, CONTRIVANCES, MANIPULATION.

Chemistry is an experimental science. Success in experiment depends on care in the preparation of apparatus, skill and cleanliness in its use, and keen observation of the conditions. The following are a few notes and hints on the simpler methods employed in the laboratory.

1. The source of Heat. The Bunsen burner. Unscrew the tube of your 'Bunsen'; you will notice a small jet attached to the stand. Connect the burner to the gas supply and set a light to the jet. A long ordinary gas flame is seen. Turn out the flame and screw on the tube. Then close the holes at the bottom of the tube and light the gas at the top, again an ordinary luminous flame appears. Open the holes and observe that the flame is now smaller, non-luminous and hotter.

Blowpipe flame. Place the nozzle of your mouth blowpipe on the edge of the tube and blow through the luminous flame (Fig. 1). It is obvious that mixing air with coal gas concentrates the flame and renders it hotter and less luminous.

Fish-tail burner. This gives a broad luminous flame suitable for softening glass tubing in order to bend it.

*Exp. Use the piece of thin iron wire supplied to find which flame and which part of the flame is the hottest. You can judge by the colour of the heated wire.

N.B.—An asterisk * prefixed to Exp. indicates a suitable experiment for the whole class to perform.

2. Modes of Heating, Evaporating and Drying.

It is evident that if we require to keep our apparatus clean we must use the smokeless flame of the Bunsen burner.

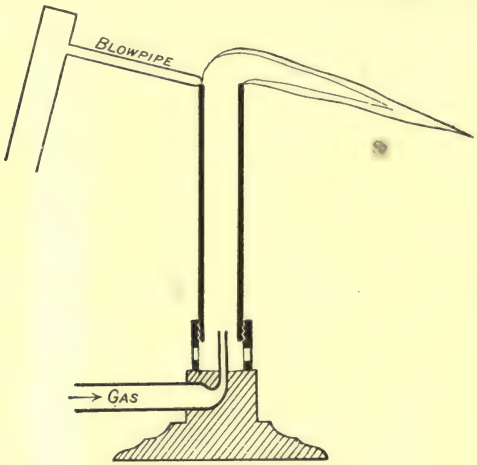


Fig. 1.

Asbestos Wire gauze is used to distribute the heat and to prevent the flame spreading round a glass vessel.

A sand bath is an iron tray filled with sand, heated below by the Bunsen flame. A steady flow of heat suitable for evaporating purposes is obtained.

A water bath is used where the temperature must not rise above 100° C. It consists of a large beaker or pan (of glass or of metal) in which water is boiled; the steam driven off heats the dish or flask placed over the boiling water (Fig. 2). It is useful when it is necessary to evaporate or boil away inflammable liquids as well as where slow evaporation is required.

A drying cone is a cylindrical or cone-shaped piece of sheet iron serrated at the top and bottom (Fig. 3). If it is placed on

a heated sand tray, a current of hot air rises through the cone. It is especially useful for drying filter paper and precipitates which might be injured by less gentle heat.

A desiccator is a glass vessel, with a large air-tight lid,

Fig. 2.

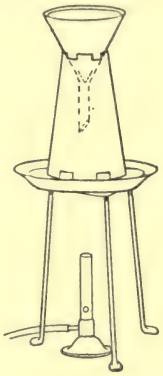


Fig. 3.

Fig. 4.

wherein a substance may be placed which might otherwise absorb moisture from the air (Fig. 4). The desiccator contains calcium chloride or strong sulphuric acid, each of which readily combines with water vapour, thus keeping the air in the desiccator dry.

3. Glass bending and blowing. To **cut** a piece of narrow glass tubing, quickly, but not heavily, draw a triangular file across the tube at the point required until a deep scratch is made. Take the tube in both hands with the thumbs meeting at the scratch and bend the thumbs outwards. If the tube does not break easily make the scratch longer with the file.

To **bend** glass tubing, use the broad luminous fish-tail burner. *Support* the tube with both hands, *rotating* it in such a way that a considerable length of tube is heated (Fig. 5). When the glass is soft it may be bent at pleasure either in

the flame or outside it. Allow the tube to cool before removing the lamp-black.

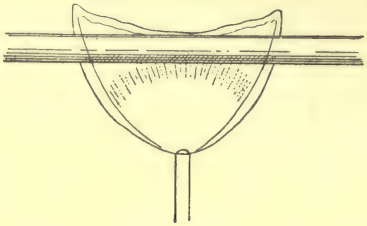


Fig. 5.

To **draw out** glass tubing (Fig. 6 A), the tube is thoroughly softened by *rotating* it in either flame, preferably the blowpipe. The hands support the tube on either side of the flame but they must not draw it out while it is still in the flame. The more rapidly the hands are drawn apart the finer the tube will be (*capillary tubing*).

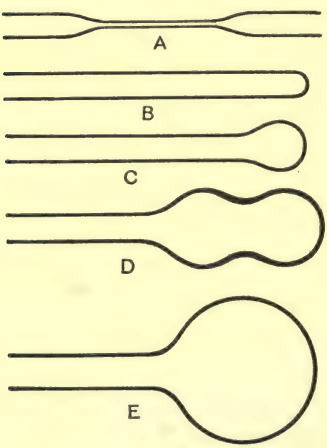


Fig. 6.

To seal a tube and blow a bulb, again *rotate* the end of the tube in the Bunsen or blowpipe flame. Its sharp edges become rounded and softened; then they gradually *close together* (Fig. 6 B). If a **bulb** is required (Fig. 6 C), continue the heating until there is a small quantity of molten glass formed. Continue to *rotate* the tube, remove it from the flame and blow down the open cold (!) end until a small bulb is formed. If a larger bulb is required (Fig. 6 D and E), the tube must be heated again and the process continued and repeated until a sufficient

quantity of molten glass is obtained in the form of a thick bulb which may then be blown out to a larger size.

4. How to heat a glass vessel containing a liquid. The *outside* of the vessel must be dry and the heat must only be applied to a part which has liquid on the other side of the glass.

To distribute the heat a piece of wire gauze or a sand bath may be used as indicated in § 2.

5. How to bore a hole in a cork. Select a cork rather larger than is required. Roll it gently under the foot until it fits the flask or tube. Select a brass or steel cork borer rather smaller than the tubing used. *Wet* the sharp end of borer and use it as if it were a gimlet, beginning with the narrow end of the cork and boring down on to a soft piece of wood. *Wet the glass tubing* before working it into the hole in the cork.

To bore a hole in a rubber cork dip the borer in *caustic potash solution* before using.

6. To test whether a piece of apparatus is air-tight, close all the openings except one which should be a narrow tube rounded so as not to cut the tongue. Suck out air at this tube and cover the end with the tongue. If the tongue remains drawn into the tube the apparatus is air-tight.

7. Filtering. Fold a round filter paper twice through the centre so that it makes first a half-circle, and secondly a quadrant. Then take three thicknesses of the paper and press inwards with the fingers so that a cone is obtained. Fit this inside the filter funnel, wet the paper and press it to the glass inside the funnel (Fig. 7).

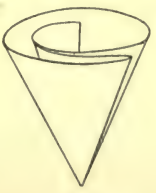


Fig. 7.

To **filter rapidly,** the filter funnel is corked into a flask with a side tube and the latter is connected with a vacuum pump.

The pressure is then slightly lessened inside the flask (Fig. 8).

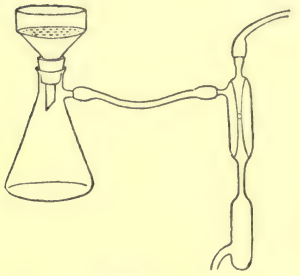


Fig. 8.

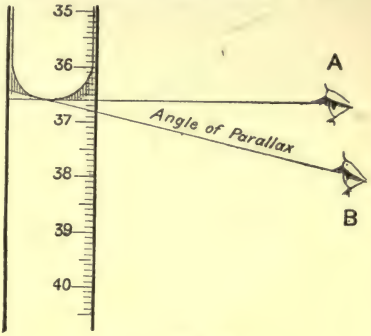


Fig. 9.

8. Parallax error. Care must be taken in reading scales, burettes, measuring cylinders, &c., to see that the line of sight from the eye to the scale is always at *right angles* to the scale. It is evident from Fig. 9 that position *A* is correct, while position *B* is not. The reading from *A* is 36·6, the reading from *B* is 36·8, the difference is called the *parallax error*

To collect a gas (*a*) by means of a pneumatic trough, we use gas jars, which, after being filled with water, are closed with ground glass plates. They are then placed mouth downwards in the trough, and the plates are removed under water. The delivery tube is passed under a '*beehive*' *shelf* and the jar is placed on the 'beehive' (see Fig. 10).

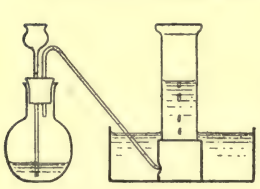


Fig. 10.

(*b*) **by means of an aspirator.** An *aspirator* is generally made by fitting a two-holed cork to a 'Winchester Quart' bottle. Through the cork a short tube and a siphon tube are put, with rubber tubes and spring taps fitted to each (see Fig. 11). Test whether the apparatus is air-tight. Fill the aspirator with water

and then by suction at the shorter tube fill the siphon tube from the beaker of water and close the tap on the siphon tube *only*.

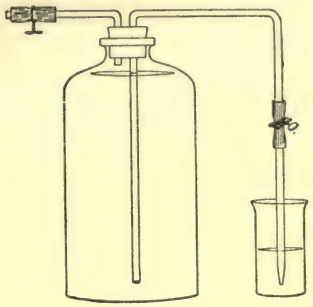


Fig. 11.

To measure the volume of gas, connect the shorter tube of the aspirator to the delivery tube of the apparatus generating the gas. Put a clean measuring cylinder under the siphon tube and open the tap. Proceed to pass the gas, and keep the end of the siphon tube under the surface of the displaced water. When reaction has ceased and the apparatus is at the temperature of the room, adjust the level of the water in the cylinder to that in the aspirator and close the tap. The volume of the water collected now equals the volume of the gas evolved.

9. Corrections of Volume of Gases for Temperature and Pressure. In measuring gases, the volume under standard conditions of temperature and pressure (S. T. P.) must be obtained, viz. at 0° C. and 760 mm. of mercury pressure. [For explanation see Part I. Physics, §§ 62, 119.]

If V is the volume at $t°$ C. and P mm. then the volume at 0° C. and 760 mm. $= \dfrac{V \times 273 \times P}{(273 + t) \times 760}$.

E.g. 100 c.c. is volume of gas collected at 15° C. and 740 mm. then the volume at 0° C. and 760 mm., *i.e.* at S. T. P.,

$$= \frac{100 \times 273 \times 740}{(273 + 15) \times 760} \text{ c.c.} = 92 \text{ c.c.}$$

Correction for Water Vapour Pressure.

If the gas is collected over water at $t°$ C. the pressure of aqueous vapour at $t°$ C. must be subtracted from the pressure of the atmosphere P. Thus in the above example, the pressure of aqueous (water) vapour at 15° C. = 12 mm., therefore the pressure due to the gas alone = 740 − 12 = 728 mm.

10. Use of Spherical Flasks. Whenever an experiment is done at pressure greatly above or below that of the atmosphere, as for instance in § 42, a round bottomed spherical flask should be used, otherwise the glass may burst or collapse.

11. To introduce a powder into a combustion tube,

Fig. 12.

fold a piece of paper once so as to form a V-shaped spill. Put the powder into the V of the spill and carefully push it into the tube, held horizontally. Turn the spill so that the powder is transferred to the tube and remove the spill (Fig. 12).

12. Rules to be observed in using the Balance.

1. See that the weights are correct before and after use and that they are in their right places in the box. The weights should be multiples and submultiples of 5, 2, 2, 1 grams or 5, 2, 1, 1.

2. Never put on or remove anything from the pans when the balance is free to swing and never leave it swinging.

3. First see that the balance is correctly adjusted and then always put the substance to be weighed in the left-hand pan.

4. Never touch the balance or the weights with your fingers if you can possibly help it.

5. Always begin with a weight that is too great and continue to add weights in the same order, downwards.

6. Never weigh anything hot or wet or likely to stick to the pan.

7. With the left hand, raise the lever gently, and only turn it completely over when a balance is nearly obtained.

8. Count the weights *both* by the number on the pan *and* by the vacant places in the box.

13. Weighing by difference.

(*a*) An empty beaker or dish may be first weighed and afterwards the substance added and the two weighed together. Then the weight of the substance is obtained 'by difference.'

(*b*) A substance may be first weighed in a test tube and afterwards some of it may be tipped into a suitable vessel. Then the tube and remainder are again weighed and the weight of what has been tipped out is found 'by difference.'

14. To enter notes on observations and results. Notes should be made as near as possible to the time of observation. Entries of numerical results should each occupy a single line, the particular unit of measurement (grams, c.c., &c.) should be added and sufficient detail should be given to prevent confusion of ideas.

Practical Exercises.

With the flasks, corks and tubing provided, fit up the following pieces of apparatus: (1) a gas generating flask (Fig. 12 A), (2) a wash bottle (Fig. 12 B).

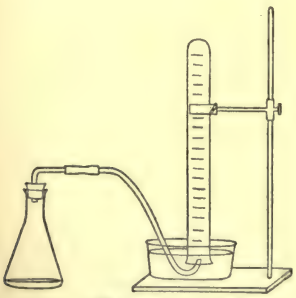


Fig. 12 A.

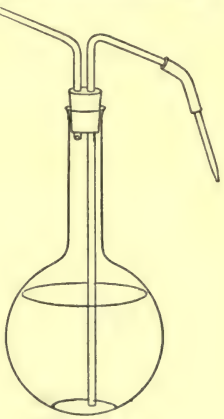


Fig. 12 B.

Questions on Chapter I.

1. Give reasons why it is necessary to attend to each of the Rules for Weighing in § 12.

2. Why is it necessary to avoid 'Parallax Error'?

3. Draw up a set of rules with a view to keeping the laboratory and apparatus clean and tidy.

CHAPTER II.

PHYSICAL AND CHEMICAL CHANGES.

15. What is Chemistry?

We know that as coal burns away in the grate quite different substances make their appearance: smoke, soot and inflammable gases come away and ashes are left: besides, heat energy is set free which may be used for various purposes. We know that we need food; that, although we are constantly eating to keep up the strength and the warmth of our bodies, the weighing machine and the tightness of our clothes tell us that not the whole of our food has been utilized in putting on weight. We know too that food may putrefy and drinks ferment. In short we are well acquainted with instances of building up, of change and of decay. Chemistry helps us to understand these processes.

Exp. Float a small candle on a cork in a basin of water (Fig. 13). Light the candle. It burns away. Put a glass jar over it. Its flame becomes more dim, flickers and goes out. Notice all that has happened. Remove the jar. Our eyes and noses tell us that both the air and the wax have taken part in the burning and that other things are formed as the candle disappears. To make clear the inner meaning of these changes is the work of the chemist.

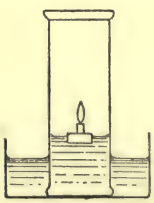


Fig. 13.

When we study **Chemistry,** we try to find out the composition of substances—how they are

built up—what changes take place when they interact—what are the conditions, and what is formed as a result.

16. Physical changes. We have already noticed that there are very complicated changes going on around us, as, for instance, the conversion of food into flesh, bones, skin and hair. But, are all changes as complicated and is it possible to classify them?

(i) *Water.* Ponds and puddles dry up. The clouds and mist form out of a clear sky. Down come the rain, hail and snow. The pond freezes and we skate on the floating ice. The ice melts and once more we see the water of the pond.

(ii) ***Exp.** Weigh a piece of *platinum*, then hold it in the Bunsen flame —it glows. Remove it from the flame—it cools and appears to be unaltered Its weight has not changed.

(iii) **Exp.** Repeat, substituting a lump of freshly burned quicklime, and use a blowpipe-flame as the source of heat. Again no permanent change can be detected.

(iv) ***Exp.** These two steel pens do not attract each other. Weigh them. Then, taking them together by the points, rub them several times in one direction only along a bar magnet. Try whether their ends attract or repel each other. The steel is now magnetized. Heat the pens red-hot. They have lost their magnetism, although they are still steel pens and their weight has not altered.

All these are instances of *Physical change* which is accompanied by:

(*a*) a change in state which may be reversed,

(*b*) no formation of new bodies,

(*c*) no alteration in composition,

(*d*) no alteration in weight.

17. Water in its three states.

***Exp.** To illustrate Physical Change use the apparatus shown in Fig. 14 to *distil* a liquid. Water is boiled in the flask *W* and the steam

is *condensed* in the flask *C*. By a thermometer *T*, we can read the temperature of both water and steam. Taste the *distillate D* and *evaporate* a little to dryness on a watch glass. Notice there is no *residue*. Freeze some of the *distilled* water in a test tube by surrounding the tube with a mixture of

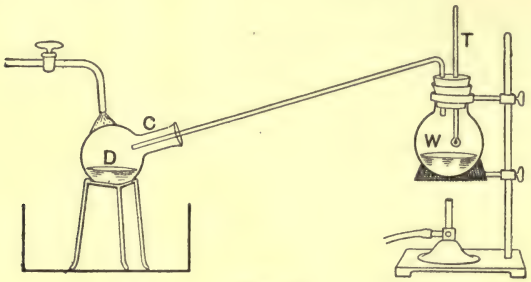


Fig. 14.

ice and salt. Warm the ice on a watch glass and evaporate the water. Again nothing remains. Describe your experiment to show the physical change of water into its *three states—solid, liquid, gas.*

[*Liebig's Condenser* is more efficient than your apparatus. The connecting tube is surrounded by a jacket of cold water. Fig. 15.]

Fig. 15.

18. Further Experiments to illustrate physical change.

***Exp.** i. Heat very *gradually* in test tubes small separate quantities of (*a*) Sulphur, (*b*) Iodine, (*c*) Mercury. Observe and record all the changes both on heating and on cooling.

***Exp.** ii. Rub *Ebonite* with *Flannel* and notice the attraction between the *electrified* ebonite and light objects such as paper. Notice there is no change in weight of both ebonite and flannel.

Exp. iii. Set a *gyroscope* rotating rapidly. Find in what directions more force has to be applied to move the whole toy when the wheel is rotating. Also find whether its weight varies in its states of motion and of rest by suspending it to a spring balance.

19. Chemical Changes. Let us first give a few instances.

(i) When we strike a match, we can see, hear, feel, taste and smell that a change is going on: but it is quite another story from that related in physical change. It reminds us of an old nursery rhyme: all the forces in creation cannot restore that match to its original condition.

(ii) **Exp.** Burn some *magnesium* ribbon: again there is a complete change accompanied by light and heat: the white ash formed does not resemble the metal magnesium.

(iii) Iron rusts, and rust is quite different from iron.

It was shown in our first experiment (§ 15) that air was necessary in burning a candle. We shall find later that the match, the magnesium and the iron, as they burn or rust, *gain* in *weight* from the air and that the air *loses* an equal weight.

Chemical changes then involve:

(*a*) the formation of new substances,

(*b*) difficulty in reversing the process,

(*c*) an alteration in temperature, and

(*d*) an alteration in weight of each substance taking part.

20. *Experiments to illustrate Chemical Change.

(i) Add water to the lump of *Quicklime* used in § 16, iii.

(ii) Mix together a little *Iodine* and *Phosphorus* (Red) in a dish.

(iii) Heat strongly *Iron* or *Copper* turnings with *Sulphur* in a test tube.

(iv) Heat on porcelain, with the aid of your blowpipe, small quantities of *Copper* (foil), *Iron* (wire), *Zinc* and *Lead*.

(v) Warm a solution of *Copper Sulphate* (Blue Vitriol) and add *Iron* turnings.

Describe carefully all the changes and examine the products.

Questions on Chapter II.

1. What are the characteristic features of (*a*) a Physical Change, (*b*) a Chemical Change?

2. Heat is applied to the following substances in air (*a*) Ice, (*b*) Sulphur, (*c*) Charcoal, (*d*) Chalk, (*e*) Sand.

Classify as chemical or physical respectively the various changes which occur, stating fully your reasons. (C. J.)

3. Describe what takes place when the following substances are heated in air:—(*a*) Mercury, (*b*) Magnesium, (*c*) Iodine.

4. Contrast the differences (*a*) on heating Iron wire with Sulphur and (*b*) boiling water in a platinum vessel.

CHAPTER III.

MIXTURES, COMPOUNDS, ELEMENTS.

21. Mixtures. The sand on the sea-shore is a mixture. You have probably washed a bucketful gradually away in a stream of water, stirring the sand so that the lighter things are carried off first. You remember that you found promiscuously crabs, shrimps, seaweed, coal, sand (fine and coarse), shells and pebbles. So in Chemistry a mixture has no definite composition, it is not the same throughout, and it is possible to separate the constituent parts by mechanical means, or by treating it in such a way that one part is affected while the rest remains unaltered.

Here is a mixture of *salt* and *beads*. An ounce from the top does not necessarily contain the same weight of beads as an ounce from the bottom of the mixture. To test this by experiment would be easy if we got rid of the salt. To pick out the beads one by one would be a tedious task. Water will dissolve the salt and leave the beads.

*Exp. Weigh out about 5 grams of the mixture in a dish. Let a slow stream of water pass over the beads and carry away the salt in solution. Soon the water does not taste of salt. *Decant* off the water and dry the beads on filter paper, finally heating them over the drying cone. Weigh the beads and calculate the percentage composition of the mixture. Although no beads have been lost it is probable that no two results are alike.

In a later chapter we shall learn how to *separate* by means of *solution* in a more accurate way.

22. Compounds. If some of the solution of salt is evaporated slowly, we find by aid of a magnifying glass that each grain is bounded by rectangular faces and is, in fact, a cubical *crystal*. All the grains taste alike and seem to have the same composition throughout (*homogeneous*). We cannot split up salt easily. If it is melted and a current of electricity is passed through the *fused* salt, a metal, *sodium*, and a gas, *chlorine*, are liberated in the proportion of 23 parts by weight of sodium to 35 parts of chlorine. It does not matter whether the salt came from Cheshire, Stassfurt or the Indian Ocean, as long as it is pure salt, it always has the same composition. If we burn *sodium* metal in *chlorine* gas, the two combine in the proportions named and pure salt (sodium chloride) is formed. This *constancy* of *composition* is the most important feature of a **chemical compound,** such as sodium chloride.

23. Elements. Sodium and chlorine have never been split up into anything simpler than sodium and chlorine—they are in fact *elements*, *i.e.* simple bodies.

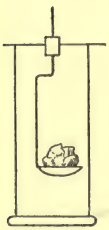


Fig. 16.

Exp. Here is a jar of the yellow, disagreeable gas, *chlorine*. Place a piece of *sodium* metal in the *deflagrating* spoon. [**Danger.** Sodium must be kept away from *water*: the two react violently.] Heat the sodium strongly and lower it into the chlorine. Combination, in the definite proportions just mentioned, between the two elements takes place, accompanied by heat, and the white salt sodium chloride—a chemical compound—is formed.

Note. This result may be shown in a striking and beautiful experiment as follows. Pass a current of chlorine, prepared by slowly dropping concentrated hydrochloric acid on crystals of potassium permanganate (see Fig. 17), over gently heated sodium in a hard glass tube. (A T tube dipping into mercury must be inserted as a safety valve.) The sodium burns brilliantly and salt (sodium chloride) is formed. The excess of chlorine should pass into the ventilating shaft or be absorbed in potash solution (use a wide funnel).

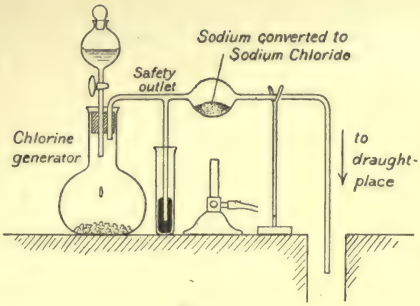


Fig. 17.

24. Properties of a compound. The following conditions accompany the formation of a chemical compound:

(*a*) It is formed by combination of elements in fixed and definite proportions by weight.

(*b*) It is impossible either to see the elements in the compound or to separate them without a chemical change.

(*c*) The compound is not like the elements of which it is composed.

(*d*) A change of temperature accompanies its formation. Contrast these with the corresponding properties of mixtures. Then refer to your experiments on chemical change, § 20, some of which may be explained further by aid of the following equations:

Elements	*Compounds*
Phosphorus combines with Iodine to form	Phosphorus Iodide
31 parts by weight + (3 × 127) parts	= 412 parts by weight
Iron combines with Sulphur to form	Iron Sulphide
56 parts by weight + 32 parts	= 88 parts by weight
Copper combines with Sulphur to form	Copper Sulphide
63 parts by weight + 32 parts	= 95 parts by weight

You must not imagine that all compounds are so easily made by *direct* union of their elements. These equations have been

given here simply to draw your attention to the fact that a particular chemical compound is always composed of the same elements united together in definite proportions by weight.

Questions on Chapter III.

1. What is an element? Name five elements and five compounds.

2. What are the characteristic properties of compounds?

3. You are asked to find out whether an earthy powder which has been found underground is a definite chemical compound or not. How would you proceed? State what you would expect each of your experiments to prove. (O. J.)

4. What is the difference between a chemical compound and a mechanical mixture?

5. State how you would (*a*) decompose salt into its elements, (*b*) obtain salt *from* its elements.

CHAPTER IV.

SEPARATION AND PURIFYING BY DISSOLVING, FILTERING, CRYSTALLIZING AND DISTILLING.

25. Separation by dissolving. We have already separated **salt** from **glass** (Exp. § 21). If we examine **gunpowder** under the microscope, we can see that it is a *mixture*, not a *compound*. Gunpowder contains (1) *nitre* (saltpetre)—white; (2) *sulphur*—yellow; (3) *charcoal*—black.

Nitre is soluble in water, the other two are not.

Sulphur is soluble in carbon bi-sulphide, charcoal is not.

Exp. To isolate the Nitre we must add *water* to the gunpowder, in a beaker. After warming and stirring the solution, the nitre will dissolve. Filtering will give us a clear solution of nitre which may be evaporated until crystals form on cooling.

To isolate the Sulphur. The filter contains sulphur and charcoal moistened with nitre solution. First pass hot water through it to remove the nitre. Then dry the filter and its contents at a temperature below 114° C. (m.p. of sulphur). Transfer the dry mixture of sulphur and charcoal to a beaker containing carbon bi-sulphide. [**Danger.** Carbon bi-sulphide is very inflammable.] Warm the contents of the beaker by placing it in a larger vessel containing hot water and stir well in the *draught-place*. Filter the hot solution of *Sulphur* in *Carbon bi-sulphide* into a shallow dish and allow it to evaporate slowly. Octahedral crystals of sulphur will form in solution (Fig. 18). Meanwhile *isolate* the *Charcoal* by washing the filter with more carbon bi-sulphide. Leave the filter to 'dry' in the draught-place.

Fig. 18.

***Exp. To find the percentage composition of a mixture of chalk and potassium chromate.** Into a *weighed* beaker, about 2 grams of the mixture are placed. *Weigh* the mixture in the beaker. Then add about 20 c.c. of water (distilled); warm to dissolve the yellow chromate. The chalk is not dissolved. Balance *two* filter papers and prepare a filter by folding them together so that the chalk will be caught on *one* of the papers. Filter the mixture into a *weighed* evaporating dish. Let the filter empty itself. Using as little water as possible, rinse the beaker so that all the chalk is washed on to the filter paper. Again allow the filter to empty itself. Pour a few drops of hot water down the sides of the filter until the liquid coming through is no longer yellow. (1) Carefully evaporate the liquid in the dish to dryness and find the *weight* of the *potassium chromate.* (2) Dry the counterpoised filter papers on the cone, separate them, place one in each scale pan and *weigh the chalk.* Calculate the percentage composition of the mixture.

Further ***Exps.** Find the percentage composition of the following mixtures:

(i) Sand and salt,

(ii) Nitre and alumina.

26. Terms and Definitions.

Solvent: Water is a solvent for nitre; carbon bi-sulphide for sulphur.

Saturated solution at a given temperature means that the solvent can dissolve no more of the substance *at that temperature.* A saturated solution, on being cooled, deposits crystals.

Crystal: a solid of regular, definite shape, enclosed by plane surfaces, and usually formed in solution on cooling.

Filtrate: the liquid which has passed through the filter.

Residue: the solid remaining on the filter paper.

Solubility of a substance *at a given temperature* is measured by the number of grams of the substance which 100 grams of water can dissolve *at that temperature.*

Mother liquor is the saturated solution from which the crystals have separated out.

27. Solubility and Solubility Curves.

***Exp.** **To find the solubility of Nitre at the temperature of the room.** Stir or shake water with *excess* of nitre (*i.e.* with more nitre than the water is capable of dissolving). In about 10 minutes a *saturated* solution will be made. Allow to stand and meanwhile weigh a porcelain evaporating dish. Take the temperature of the solution. Decant about 20 c.c. of the clear saturated solution into the dish. Weigh again. Then evaporate to dryness on the sand bath, but beware of over-heating. Weigh when cold. The loss of weight is the weight of water which can dissolve the amount of nitre remaining in the dish after evaporation. Heat again very gently and weigh. Repeat until the weight is constant. Calculate the weight of nitre which 100 grams of water can dissolve, at the specified temperature.

By making saturated solutions at different temperatures it is possible to obtain the solubility by somewhat similar experiments. Results may be tabulated as follows:

Nitre.	At 0° C.	100 grams of water		dissolved	13 grams
	20° C.	„	„	„	31 „
	40° C.	„	„	„	64 „
	70° C.	„	„	„	132 „

Hence, by aid of squared paper, we can plot

$$\frac{\text{weight of salt dissolved in 100 grams of water}}{\text{temperature}} = \frac{\text{solubility}}{\text{temperature}}$$

and obtain a *curve of solubility* (Fig. 19).

28. Separation by Difference of Solubility. (Fractional Crystallization.)

Exp. To separate nitre and chlorate of potash (both soluble in water). From the curves we read:

Solubility of nitre at 15° C. = 26.

Solubility of chlorate of potash at 15° C. = 5.

It is fairly evident that a solution containing these two substances will first yield crystals of the less soluble chlorate of potash on evaporating and cooling and that the *mother liquor* will chiefly contain the more soluble nitre. Re-dissolve the first crop of crystals and again evaporate and crystallize, thus obtaining purer chlorate of potash. In this way the less soluble substance is removed from solution. The salts of Radium have been isolated by this method.

Exp. Similarly separate (*a*) blue vitriol from chlorate of potash, (*b*) salt from alum.

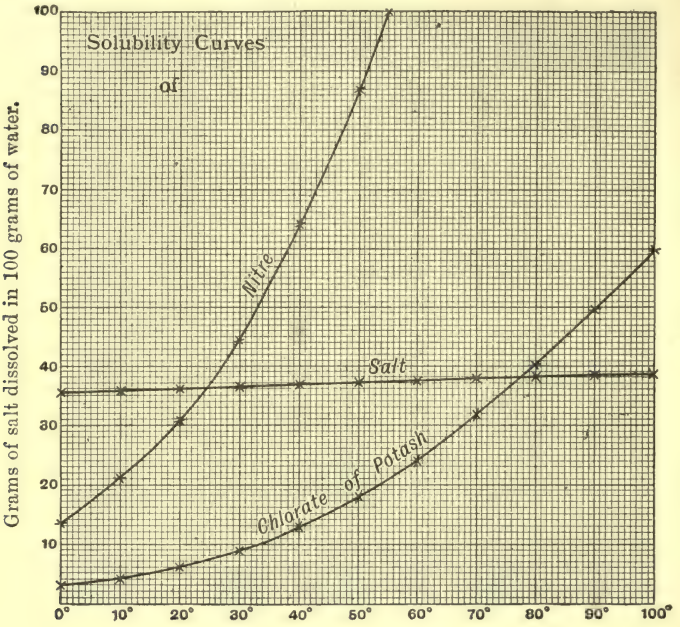


Fig. 19[1].

29. Separation of mixtures of liquids having different boiling points. (Fractional distillation.)

Ether, b.p. 35° C.; Alcohol, b.p. 78° C.; Water, b.p. 100° C. Take for example a mixture of these three liquids and distil them, using Liebig's condenser (see Fig. 15). Ether begins to distil over at about 35° C. and nearly the whole of it will be collected before the temperature has risen to 40° C., but as the other two

[1] Reproduced by kind permission of Messrs Macmillan and Co.

liquids exert vapour pressure a certain amount of their vapours will be carried over at the same time. If we collect the distillate in fractions in 3 flasks at (say) 40°, 85° and 100°, we shall have separated the liquids to a great extent. If we now take each fraction separately and re-distil it, smaller fractions will be obtained, which may be again distilled, and we can collect in one flask the liquid (say ether) coming over at 35° C. which will be practically pure.

*__Exp. Distil a mixture of alcohol and water,__ heating the distilling flask in a water bath. Collect the fraction, b.p. 78°—84°. Re-distil this separately and collect the fraction, b.p. 78°—80°. Finally distil this, adding a lump of quicklime (§ 20 (i)), when the alcohol will come over pure at 78° C.

Questions on Chapter IV.

1. What is distillation? For what general purposes is the process used in practical chemistry? Describe carefully any process of distillation which you have performed or seen in the laboratory and make a sketch of the apparatus used. (o. J.)

2. How would you ascertain whether plaster of Paris is soluble in water?

3. Suggest methods for separating (*a*) sulphur and charcoal, (*b*) water and acetone (a liquid which boils at 55° C.), (*c*) salt and chlorate of potash.

4. Calculate percentage of (*a*) sand and nitre, (*b*) sand and potassium chromate present in a mixture from the following:

	(*a*)	(*b*)
Weight of filter paper	= ·895 gm.	·78 gm.
,, ,, ,, and mixture	= 4·275 gm.	3·52 gm.
,, ,, ,, ,, dry sand	= 1·425 gm.	3·27 gm.
,, evaporating dish	= 27·35 gm.	49·75 gm.
,, ,, ,, and dry salt	= 30·2 gm.	50·0 gm.

5. What do you mean by a 'saturated solution'? How would you prepare a saturated solution of common salt and determine the 'solubility' of salt?

6. 6·8 gm. of a saturated solution of nitre yield 1·8 gm. of dry nitre. (*a*) Calculate the solubility. (*b*) What is the temperature of saturation?

7. 24 c.c. of a solution weigh 26·65 gm. and yield 7·09 gm. of solid on evaporation. Find (*a*) amount of solid in 100 c.c. of solution, (*b*) the 'solubility' of the solid.

8. Calculate the solubility of nitre from the following:

Weight of evaporating dish empty = 44·32 gm.
" " " and solution = 48·97 gm.
" " " " dry nitre = 45·14 gm.

9. If 100 grams of a solution of chlorate of potash saturated at 70° C. be cooled to 16° C. how much salt will separate out?

10. Draw solubility curves for alum and potassium chloride from the following data:

Temperature	0°	10°	20°	30°	40°	50°	60°	70°
Solubility of alum	3·9	9·5	15·1	22·0	30·9	44·1	66·6	90·7
Solubility of potass. chlor.	28	31·5	35	38·5	41·7	45	48·5	52

Find from these (*a*) at what temperature equal weights of the two substances will be dissolved by 100 gm. of water.

(*b*) How much water will be required to dissolve 17 gm. of alum at 42° C.

(*c*) What weight of potassium chloride will dissolve in 50 gm. of water at 36° C.

CHAPTER V.

SOLUTION MAY BE ACCOMPANIED BY CHEMICAL CHANGE.

30. When we make effervescing drinks, as for instance when we pour water on to sherbet, fruit salts or Seidlitz powders, we meet with instances of chemical change accompanying solution. The fact that a gas is set free gives us sufficient proof of this, for it is evident that an entirely different substance is produced when solution takes place.

***Exp.** i. Make up a mixture of *chalk* and *water* and add a few drops of hydrochloric acid. The solid chalk suspended in the water is dissolved and a gas (carbon di-oxide) is set free. This gas puts out a burning taper.

***Exp.** ii. Drop thin shavings of *zinc* into dilute sulphuric acid (oil of vitriol) in a test tube fitted with a cork and delivery tube. The *zinc dissolves* and an inflammable gas (hydrogen) is liberated, which may be collected by water displacement if desired (Fig. 20).

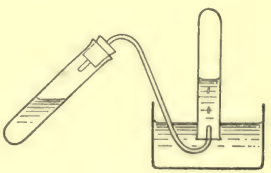


Fig. 20.

Exp. iii. Repeat the experiment, substituting (*a*) iron and (*b*) magnesium for zinc.

Exp. iv. Drop aluminium foil into a warm solution of caustic soda. Hydrogen is liberated.

In all these cases it is evident that solution is accompanied by chemical change.

31. Quantitative nature of Chemical Change.

There is a **definite relation between the weight of zinc dissolved, the weight of acid required and the weight of gas liberated.**

***Exp.** i. Weigh out exactly a *gram of zinc* foil, and put it into a small tube. Having placed about 50 c.c. of dilute *sulphuric* acid in a small flask, suspend the tube containing zinc in the flask by a thread and connect the

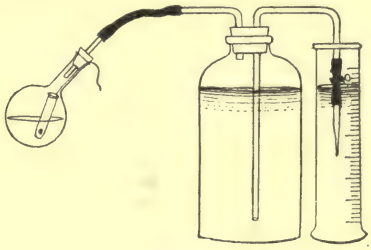


Fig. 21.

flask to an aspirator (Fig. 21). Adjust the aspirator and attend to the directions given in § 8. Tilt the flask so that the zinc is attacked by the acid and *measure the volume* of hydrogen gas liberated at the temperature and pressure of the room. If care is taken all the members of the class will obtain, within one per cent., the same volume and therefore *the same weight* of gas. (See Physics—weight of 1 litre of gas.)

It is difficult to show accurately that

(*a*) **the same weights of acid dissolve equal weights of zinc,**

and that (*b*) **corresponding weights of substance are formed in solution.**

However, approximate results are obtained as follows:

Fig. 22.

***Exp.** ii. Weigh separately two pieces of zinc rod. Suspend each of them in separate quantities of 10 c.c. of dilute sulphuric acid (Fig. 22). After some hours no more gas is given off.

(A) Carefully remove, wash and dry the pieces of zinc. Each piece will have lost the same amount.

(B) Evaporate the two solutions to an *equal* bulk and let them stand for a few hours. Collect the crystals (zinc sulphate) and dry them separately. Approximately equal weights are obtained.

These experiments repeatedly call attention to the fact that chemical reactions do not go on in a haphazard way. *The weights of substances taking part in the various changes bear a definite relation to each other* and certain conditions bring corresponding results.

Questions on Chapter V.

1. Salt 'dissolves' in water; zinc 'dissolves' in dilute sulphuric acid. Compare the two processes. How would you prove that in the second case a definite amount of sulphuric acid is required for the solution of a definite amount of zinc? (o. j.)

2. When 1 gm. of magnesium is acted upon by sulphuric acid 960 c.c. of hydrogen are collected. ·25 gm. of magnesium yields 240 c.c. of gas. What evidence is there that a chemical change has taken place?

3. If a mixture of sand, salt and chalk were given you how would you proceed in order to obtain from it all the sand and all the salt which it contains? (o. j.)

4. 10 c.c. and 20 c.c. of acid of the same strength are placed in separate beakers and the same weight of magnesium ribbon is placed in each. When all chemical action is over the metal remaining in each beaker is dried and weighed. What will be the ratio of the weights of magnesium dissolved in each quantity of acid? What proportion of magnesium sulphate will be present in each solution?

CHAPTER VI.

THE ATMOSPHERE AND COMBUSTION IN AIR.

32. Air is a mixture. We have learnt in our work in Physics that the air has weight, that it is elastic (*i.e.* if compressed, it shrinks, but expands when the pressure is relaxed) and that its volume is affected by heat and cold.

Let us now consider air from a chemical point of view. Is it a *single* substance, elementary or compound, or is it a *mixture*?

(1) Compare *air* with *steam*. Steam is a colourless gas; if we cool it or compress it, water is formed; if we cool the water further ice is formed. Similarly it is possible to cool and compress air until at a very low temperature a clear liquid is obtained which turns into a white solid on further cooling. If air is a single substance it will have a constant boiling point, just as we have found alcohol and water to have. But liquid air has not a definite boiling point, and if we collect the gas that boils away first, at the lower temperature (– 196° C.), it is quite different in density and chemical properties from the gas which comes off as the boiling point rises (– 183° C.).

The liquid with the lower boiling point is NITROGEN.

„ „ „ higher „ „ OXYGEN.

These two substances have never been divided into anything else. They are elements, mixed together in the air. Other gases are present in small quantities, the most important of

which is water vapour, but for the present we shall confine our attention to *nitrogen* and *oxygen*.

(2) If we take these gases in the right proportions and mix them together, there is no change in volume, no heat is liberated, and air is formed.

We have now two sufficient reasons for considering air to be a *mixture*, not a compound. But we shall find that there are other reasons besides.

33. Discovery of Oxygen. Oxygen was discovered accidentally by Priestley, an Englishman, in 1774. Having filled a glass jar with mercury in a pneumatic trough, he put the substance on which he wished to experiment inside the jar, and there, as it floated on the mercury, he heated it with the rays of the sun concentrated by a powerful 'burning-glass.' He heated the *calx* or ash of *mercury* (red precipitate) among the rest, and obtained a colourless gas, in which a candle burned brilliantly; and some mice, and Priestley himself, on breathing this 'air' seemed to experience a peculiar liveliness.

34. Lavoisier's work on combustion. The discovery made by two Oxford chemists that the products of burning a substance in air weigh more than the original substance had arrested the attention of a French chemist, **Lavoisier.** In 1772 he had shown that in converting *tin* into its calx (*i.e.* when tin is burnt) an absorption of air takes place. On being shown the result of Priestley's discovery and remembering that calx of mercury is formed by heating mercury in air, Lavoisier argued that the gain in weight on heating (burning) tin and mercury had been caused by the air combining with the metal. He proceeded, in 1775, to verify his supposition by the following experiments:

I. Mercury was heated (300° C.) in a closed vessel connected with a pneumatic trough (Fig. 23). Red calx formed and the volume of air diminished until after heating for several days there was no further shrinkage (= about $\frac{1}{5}$).

II. The calx was collected and heated more strongly (360° C.), when it yielded a gas whose volume equalled the

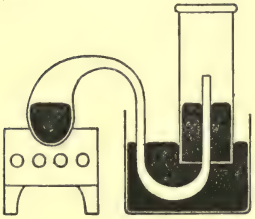


Fig. 23.

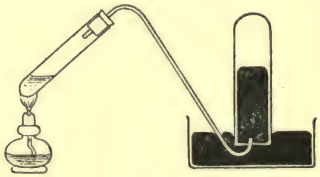


Fig. 24.

shrinkage in the first experiment (Fig. 24). This gas was added to the residual air of the first experiment, when ordinary air was again produced. Moreover the mercury obtained by heating the calx was added to the residue of mercury which was first heated with the result that the total weight of mercury was the same as at the beginning.

Lavoisier first gave the true explanation of burning, viz. that the combustible body combined with the active part of the air. He also proved not only that the burned body is not destroyed, but also that matter in general is indestructible. This second discovery may be stated as follows: *The total weight of matter taking part in a chemical change remains unaltered throughout the various processes.* **Matter can neither be created nor destroyed.** Lavoisier afterwards named the active one-fifth of the air OXYGEN and the remaining four-fifths **azote** (lifeless, inert); we call it NITROGEN.

***Exp. Effect of Heat on the Red Oxide of Mercury.** Weigh out about two grams of the red calx of mercury in a test tube. Heat strongly; put a glowing splinter into the gas given off. Continue heating until only mercury remains. Weigh again and note the loss of weight. Calculate the percentage composition of the red oxide of mercury.

35. Magnesium, Copper and Lead gain weight in burning.

***Exp.** i. **Effect of Heat on Magnesium.** Weigh a crucible and lid. Place in it about a gram of magnesium turnings or ribbon. Weigh again. Heat, raising the lid slightly from time to time to admit air, until all the metal is burned. Weigh, when the crucible is cold, and calculate the percentage *gain in weight.*

***Exp.** ii. Repeat the above experiment, substituting dry precipitated *copper* for magnesium. With the crucible weigh an iron nail for stirring the heated copper.

In a similar way we may show that *lead in burning to its yellow oxide* (litharge) *gains in weight.* The process is a lengthy one. Use a nail to stir the molten metal and expose constantly a fresh bright surface to the air.

36. Iron gains weight as it rusts.

Again *iron* may be weighed in a dish, moistened and placed in a warm place. In a few days *rust* will have formed. If now the dish and its contents are heated to expel any moisture present and are afterwards weighed, there is found to be a considerable *gain in weight.*

No rusting takes place if air is excluded. That *water alone does not cause iron to rust* may be shown by boiling together, in a round-bottomed flask for some minutes, *water* and *iron nails.* Cork or seal up the flask while the water is boiling, *i.e.* when the flask is full of steam. Another method is to pour paraffin on the surface of the water immediately after removing the source of heat. Air is thus excluded and no rusting occurs even after many days have elapsed.

[**Note.** That moisture and air alone do not cause iron to rust if carbonic acid gas is absent is proved by suspending bright iron in a closed flask containing a little potash solution, which absorbs carbonic acid gas.]

37. Oxygen from Red Lead. It is probable that in 1727 oxygen was first isolated by heating red lead (minium) but it was not recognized as a new element. In 1774, Scheele, a Swedish chemist, discovered oxygen independently of and possibly earlier than Priestley. He first obtained it by heating nitre. He also found that, by varying the conditions and the temperature, lead yields two calxes, the one yellow (litharge), the other vermilion (red lead or minium). Lead, burned at a bright red

heat, gains in weight and forms litharge. Litharge roasted at a dull red heat for some hours in a current of air gains weight still further and turns to red lead. If the temperature is raised to a bright red heat the red lead loses weight and yields oxygen, at the same time reverting to litharge.

***Exp.** Weigh a test tube containing about 10 grams of *red lead*. Heat strongly and test the gas evolved with a glowing splinter. Oxygen ('active air') is liberated and on weighing again a loss of weight is observed. Break open the test tube and notice that litharge has been formed.

38. Brin's oxygen process.

We have now learned how to use an oxide of lead to obtain oxygen from air. Baryta, an oxide of the metal barium, also gains weight on being heated in air, combining with oxygen to form a *peroxide*. At a higher temperature the process is reversed—the peroxide returns to baryta and yields oxygen. As a commercial process it is more economical to keep the temperature constant. Air is first pumped in, on the oxide heated in iron cylinders, and afterwards the *pressure* is reduced; the same changes then go on without varying the temperature. This process has been superseded by the method of obtaining oxygen from air liquefied by pressure and cooling (§ 32).

Questions on Chapter VI.

1. By what arguments can you prove that air is (*a*) not an element, (*b*) not a chemical compound? (O. J.)

2. The residues obtained by heating solid substances in air are in some cases lighter, in others heavier, than the original substances. Give one instance of each kind. Describe experiments for investigating the action in the instances you have selected and give the conclusions that may be drawn therefrom. (C. J.)

3. Describe experiments that may be made to investigate what occurs when iron rusts in moist air. (C. J.)

4. What leads us to suppose that matter is indestructible?

5. How was oxygen first obtained?

6. Calculate percentage increase in weight of magnesium on being heated in air from the following:

Weight of crucible = 14·775 gm.
" " and magnesium = 14·805 gm.
" " " magn. oxide = 14·825 gm.

7. One gram of mercuric oxide yields ·074 gm. of oxygen on being heated. If 2 gm. of oxygen are required what weight of mercuric oxide must be heated?

CHAPTER VII.

NITROGEN FROM AIR. PROBLEMS ON BURNING.

39. Nitrogen. The active part of the air (oxygen) may be removed by a process similar to Lavoisier's. Instead of heating *tin* in a closed space, *Phosphorus*, a readily inflammable substance, is used. The phosphorus burns, combining with oxygen, and leaves *Nitrogen*, mixed with small quantities of other inert gases.

Exp. Float a small dish on water in a deep pneumatic trough. A small piece of yellow phosphorus is placed on the floating dish. [**Danger.** *Phosphorus (yellow) should be kept and cut under water and never touched with the hands.*] Having put a wide stoppered glass jar over the dish, light the phosphorus by touching it with a hot wire. Immediately insert the stopper, and note carefully all that takes place. Pour water into the trough until the levels inside and outside the jar are the same, and make a rough estimate of the diminution in the volume of the air. Wait until the solid white *oxide of phosphorus* has dissolved. The remaining gas is chiefly *nitrogen*. Try the effect of putting a burning taper into it.

Fig. 25.

40. To prove that phosphorus gains in weight on burning in air and to collect the 'inactive' air (Nitrogen).

Exp. Set up the apparatus as shown in Fig. 26. One of the two aspirators, which are of equal size, is filled with water. The two aspirators are connected by their shorter tubes to a piece of combustion tubing loosely

packed at each end with asbestos wool, enclosing 2 or 3 grams of red phosphorus. We can pass the air contained in the aspirator and tube backwards and forwards over the phosphorus until it no longer burns by siphoning the water alternately from each aspirator, raising and lowering the beakers accordingly. Having weighed the tube and contents, start the current of air by opening the taps and blowing gently down the rubber tube on the left; close the end, and then put it under the water in the beaker. Ignite the phosphorus by warming the tube and then remove the flame. The white oxide of phosphorus is caught in the asbestos. After a few minutes the phosphorus goes out and will not burn again even if heat is applied. Stop the siphoning when one of the aspirators is full of water. It will then be seen that the other aspirator is no longer empty but contains

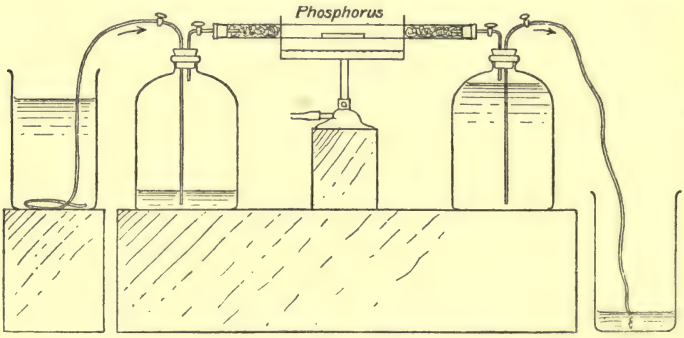


Fig. 26.

about $\frac{1}{5}$ of its volume of water which represents the volume of active air absorbed by the phosphorus, the remaining $\frac{4}{5}$ being *nitrogen*. Close the taps and keep this nitrogen for further experiments. Weigh the central tube and notice the *increase in weight*.

N.B. *Copper turnings* heated to redness may be substituted for phosphorus, but the absorption of oxygen is slower.

41. To compare rusting (of iron) and smouldering (of phosphorus) with burning.

Exp. Connect the longer tube of the aspirator with the water tap and fill two gas jars with the *nitrogen* prepared in Exp. § 40. (See Fig. 27.)

(1) (*a*) Into one of these introduce some iron filings by putting some in the middle of a ground glass plate, and covering the end of the collecting

jar under water with the glass plate. Now shake the jar about to distribute the filings over the inside of the jar and put it back again in the trough.

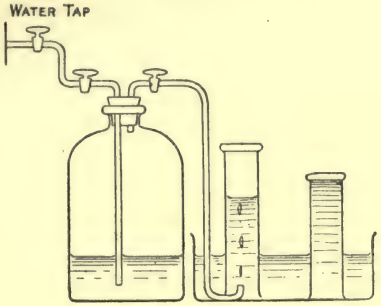


Fig. 27.

(*b*) Take another gas cylinder full of *air*; moisten the sides and sprinkle iron filings inside as before. Invert this by the side of the jar containing iron in nitrogen.

(2) Attach two pieces of yellow phosphorus to two pieces of wire [**Danger.** Do this under water] and put the phosphorus into (*c*) a jar of nitrogen, (*d*) a jar of air.

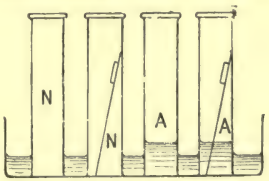


Fig. 28.

Allow these 4 cylinders to stand side by side for 2 or 3 days and notice what takes place. No change is observed in the two jars of nitrogen; but in the jars of air smouldering and rusting continue until the water has risen to fill $\frac{1}{5}$ of the jar. Cover the ends of the cylinders with glass plates and test the residual gas in each case with a burning taper. Write down your deductions from the results of this experiment, taking into consideration §§ 36 and 40.

42. The air loses as much as the burning substance gains in weight. The principle of the *indestructibility of matter* (§ 34) is here applied to the problem of burning.

Exp. *To show that the total weight of air and phosphorus remains unchanged in burning.* Cut a small piece of yellow phosphorus. [**Danger.** Cut it

under water.] Hang a round bottomed flask containing a little asbestos wool on the balance and put a fitting rubber cork in the pan. Having obtained the approximate weight, carefully take the phosphorus out of the water and dry it on filter paper. Then drop the phosphorus into the flask and on to the asbestos. Weigh accurately the whole apparatus *with the cork in the pan.* Insert the cork tightly. Warm a small part of the flask[1] and roll the asbestos gently over the heated glass until the phosphorus catches fire. [**Danger.** *The flask has been known to burst in this experiment. Directly the phosphorus catches fire cover the flask with a duster.*] Allow to cool and weigh. *No change* is noticed. Open the flask under water, 'level the water,' and insert the cork. Measurement will show that again $\frac{1}{5}$th of the air has been absorbed by the phosphorus to form oxide of phosphorus.

43. The substances produced when a candle burns weigh more than the candle.

Exp. Cut a piece of wire gauze and suspend it by wire so as to form a shelf about half-way down a long lamp chimney. On this shelf pack loosely lumps of quicklime with sticks of caustic soda above the quicklime. These substances absorb the products of combustion of a candle burning on a perforated cork fitted to the bottom of the tube. Weigh the apparatus *before* lighting the candle, and again when the candle has been burning for a few minutes. Notice that there is a *gain in weight.* (See Fig. 29.)

44. The gain in weight comes from the air.

Exp. The candle is next burned in a closed flask containing air. The flask is counterpoised on the balance and the candle ignited by means of a fine platinum wire heated to redness by passing a strong electric current as indicated in the foot-note to § 42. The fact that the weight is unaltered proves that the gain in weight observed in § 42 must come from the air and also shows that, although the candle burns away, *matter is not actually destroyed.* (See Fig. 30.)

45. To prepare pure nitrogen.

Nitrogen prepared from air by removal of oxygen contains about one per cent. of an inert gas argon and traces of

[1] Here again there is *danger.* The following is the best method. The shaft of a deflagrating spoon and a piece of copper wire pass through the rubber cork and are connected at the spoon by a piece of *fine* platinum wire. The phosphorus is placed on the spoon and touching the fine wire which is heated by passing a strong electric current. (See Fig. 30.)

several others. To prepare nitrogen free from these impurities, 15 grams of ammonium chloride (sal ammoniac) are mixed with 20 grams of sodium nitrite in a 300 c.c. flask. 100 c.c. of water are added and a corked delivery tube is fitted to the flask. On warming the mixture in the flask gently, pure nitrogen is set free and may be collected in cylinders by the pneumatic trough method.

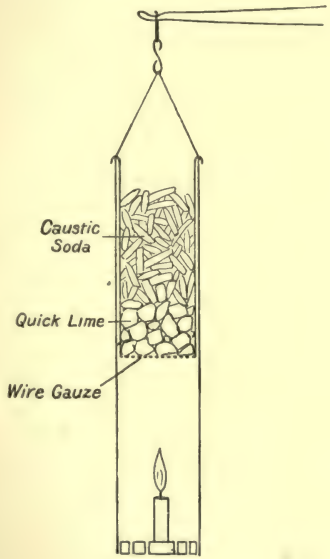


Fig. 29.

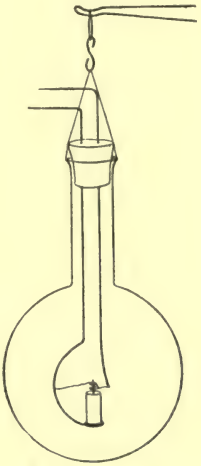


Fig. 30.

Questions on Chapter VII.

1. Describe an experiment to prove that air contains approximately four-fifths of its volume of Nitrogen.

2. How would you prove that, when a substance burns, the air loses as much as the burning substance gains in weight?

3. Some moist iron filings are placed in a jar of nitrogen and some more in a jar of air. After being left for some days a lighted taper is introduced into each jar. State and explain what happens, and whether there has been any change in the volume of the gases.

4. How is Nitrogen obtained? What experiments would you make with it in order to find out its properties? (C. J.)

5. How could you prove that only one of the two chief constituents of the atmosphere is concerned in ordinary combustion?

CHAPTER VIII.

OXYGEN.

46. We have found that Oxygen may be obtained indirectly from the air by means of the metals mercury and lead. We must consider a more convenient method of preparing the gas in the laboratory.

***Exp.** i. **Oxygen from potassium chlorate.** Heat a few crystals of *potassium chlorate* in a test tube. The crystals split and crackle (decrepitate), melt and decompose. Try whether a glowing splinter is ignited by the gas driven off by heating. The residue contains no oxygen and is called *potassium chloride.*

Potassium chlorate = potassium chloride + oxygen.

***Exp.** ii. **To find the weight and volume of oxygen obtained by heating one gram of potassium chlorate.** A hard glass test tube is *weighed* and into the bottom of the tube from ·5 to 1 gram of finely ground potassium chlorate is introduced by means of a spill of paper (§ 11). *Weigh the chlorate* accurately 'by difference.' Get ready an aspirator (§ 8) and connect the test tube to it (Fig. 31). Heat the chlorate gently and afterwards strongly until no more oxygen is set free. Allow the tube to cool. Measure the volume of water displaced, which equals the *volume of*

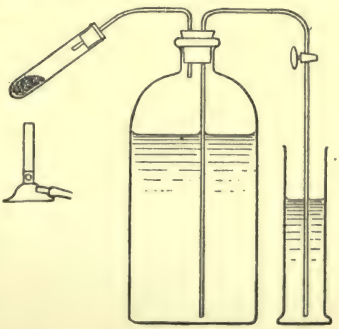


Fig. 31.

oxygen liberated. *Weigh the tube* and potassium chloride remaining. The *loss of weight* = wt. of oxygen. Note the temperature of the water and the barometric pressure. Enter your results after the following manner:

Weight of hard glass tube (empty)	= 22·84 gram
„ „ „ + potassium chlorate	= 23·56 gram
Weight of potassium chlorate used	= 0·72 gram
Vol. of water displaced = vol. of oxygen	= 216 c.c.
Temperature = 15° C. Pressure of atmosphere	= 740 mm.
Pressure of aqueous vapour at 15° C.	= 12 mm.
Weight of tube + potassium chloride (after heating)	= 23·28 gram
∴ weight of oxygen displaced	= 0·28 gram

Volume of oxygen at S.T.P. (see § 9)

$$= \frac{216 \times (740 - 12) \times 273}{760 \times 288} = 196 \text{ c.c.}$$

For the answers required, divide the last two results by 0·72.

47. Preparation of Oxygen. Considerable heat is required to decompose potassium chlorate unless a second substance is present which helps to break up the chlorate. The presence of *sand* and of *black oxide of manganese*[1] aids decomposition at a lower temperature.

***Exp.** i. Put about a gram of *potassium chlorate* in a test tube. Heat slowly so as to melt the potassium chlorate without decomposing it. Test that no oxygen is given off. Now add a few grains of black oxide of manganese; oxygen is liberated rapidly.

***Exp.** ii. Warm gently a mixture of about 8 grams of *potassium chlorate* and 2 grams of *black oxide of manganese* in a small hard glass flask connected by cork and tube to the pneumatic trough (Fig. 32) and collect several cylinders full of oxygen 'by water displacement.' The cloudiness of the gas is due to the presence of a fine white dust of potassium chloride, carried over in the rush of oxygen.

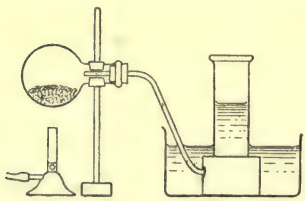


Fig. 32.

Potassium chlorate = potassium chloride + oxygen.

[1] These substances, although they may change during the process, are found to be unaltered after the reaction is completed. They are called **catalytic agents** or *catalysts*, *i.e.* 'breakers down.'

Exp.** iii. **Solubility of Oxygen.** Take a test tube or flask small enough to be closed by your thumb. Nearly fill this with oxygen. Having closed the flask with your thumb shake up the oxygen with the undisplaced water for a few minutes. Then open the flask under water and notice that a little more water enters, showing that oxygen is ***slightly soluble in water.

Fish, water-beetles, shrimps, lobsters, crabs, shell-fish and other aquatic animals, which do not need to rise to the surface for breath, use the dissolved oxygen to sustain life by keeping a constant stream of water passing over their gills or other breathing apparatus, just as we constantly take fresh air into our lungs.

In Chap. XII we shall learn how to collect and analyse the air dissolved in spring water.

48. Combustions in Oxygen. Indicators. In order to study what is produced when elements burn in oxygen, we try how the products affect a dye called **litmus,** prepared from certain lichens. Sour liquids, called **acids** (vinegar, for instance), turn vegetable dyes to a *reddish* colour. Red cabbage pickle is not crimson until vinegar, which contains acetic (= acidic) acid, is poured upon it. These dyes are used to show the presence of *acids* and are called **indicators.** The blue colour of the litmus is restored on the addition of an *alkali*, such as potash or soda.

***Experiments.** In the following table notes are given on burning seven elements in oxygen. Heat is applied to start the combustion. The first five elements named are burned in a 'deflagrating spoon' (Fig. 33).

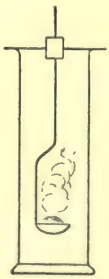


Fig. 33.

After the reaction is completed, **neutral litmus** (a solution which *reddens* on adding a drop of *acid*, but turns *blue* in the presence of a trace of *alkali*) is added to ascertain whether the oxide produced dissolves in water to form an acid or an alkali.

49. Table of Combustions in Oxygen.

	Element	Remarks	Observations	Effect on neutral litmus solution	The oxide forms with water an
non-metals	1. **Phosphorus**	Use the *red* variety for safety	Brilliant, hot flame. Phosphoric oxide formed. White soluble solid	Litmus turned *red*	Acid
non-metals	2. **Sulphur**	Use 'flour' or 'roll'	Violet flame. Sulphurous oxide is a colourless soluble pungent gas	Litmus turned *red*	Acid
non-metals	3. **Carbon**	Use a piece of dry charcoal	Sparks formed (scintillation). Carbon di-oxide is a colourless soluble gas, 'turns lime-water milky'	Litmus turned *reddish*	Weak acid
metals	4. [1]**Sodium**	**[Danger.]** A bright soft metal. Heat strongly in a *dry* 'spoon'	Golden yellow flame. Sodium oxide is a white soluble solid	Litmus turned *blue*	Alkali
metals	5. **Calcium**	Use turnings or chips of the metal: heat strongly	An orange-red flame. Calcium oxide is white and sparingly soluble	Litmus turned *blue*	Alkali
metals	6. **Magnesium**	Plunge the burning ribbon into the jar of oxygen	Brilliant light. Magnesia is a white solid oxide, slightly soluble	Litmus turned *blue*	Very weak alkali
metals	7. **Iron**	Ignite wire or watch-spring by attaching charcoal or sulphur which burn readily	Iron oxide is a greyish-blue insoluble solid	Litmus is unaffected	Insoluble in water: neutral

[1] Not to be handled by inexperienced pupils.

50. Metals and Non-metals.

Metals. We are well acquainted with *metals* such as gold, silver, iron, copper and tin; they **are bright, heavy, lustrous elements which conduct heat and electricity.** All of these named are both *malleable* and *ductile*, *i.e.* they may be *hammered* into thin sheets and also *drawn out* into fine wire.

Three of the four metals which we burned in oxygen yield oxides soluble in water. Sodium oxide combines with water to form sodium hydrate. Calcium oxide (quicklime) combines with water to form calcium hydrate (slaked lime): its solution is called lime-water. **The solutions of metallic oxides are alkaline.** The oxide of iron does not dissolve in water and therefore it does not affect litmus.

Non-metals. Phosphorus, sulphur and carbon, in their ordinary forms, do not show any of the properties of metals; they belong to the class of *non-metallic elements.* Their oxides dissolve in water to form **acids.**

The name—Oxygen. Lavoisier gave oxygen its name. By derivation it means *acid-producer.* The metals sodium, calcium and magnesium were not discovered in his day. The only soluble oxides known to him were those of the *non-metals* and these gave *acid* reactions. Hence he argued that oxygen was an essential constituent of all acids. We shall find later that there are some acids which contain no oxygen and also that oxides of metals, far from being acid-forming, can remove the acid properties from acids, whether they are soluble in water or not. (See Basic Oxides, §§ 58 and 60.)

51. Common Oxides. We must acquaint ourselves thoroughly with many of the commoner oxides for these are often recurring in our chemical work. We should examine specimens and handle them if possible.

Table of Common Oxides.

Non-metallic Elements	Oxide [anhydride = acid 'without water']	Properties of Oxide (acid forming)
Carbon: charcoal graphite diamond	(1) Carbonic oxide (carbon mon-oxide)	Colourless, poisonous gas, burns with a blue flame above red-hot coal fires
	(2) Carbon di-oxide (carbonic anhydride)	Colourless gas, 'turns lime-water milky,' forms chalk with lime
Sulphur: roll flour prismatic amorphous	(1) Sulphur di-oxide (sulphurous anhydride)	Colourless, irritating gas, forms sulphurous acid with water
	(2) Sulphur tri-oxide (sulphuric anhydride)	White needle-like crystals, forms sulphuric acid with water
Phosphorus: red yellow	Phosphoric oxide or anhydride	Has great affinity for water—a drying reagent, forms phosphoric acid with water
Silicon	Silica, sand, quartz	Is a constituent of many rocks

Metallic Elements	Oxide	Properties of Oxide Basic (see § 60). Counteract acids. If soluble, they produce *alkalies*
Sodium	Sodium oxide	Dissolves readily in water, forming sodium hydroxide or hydrate (caustic soda)
Potassium	Potassium oxide	Dissolves readily in water, forming potassium hydroxide or hydrate (caustic potash)
Calcium	Calcium oxide or quicklime	Forms slaked lime (calcium hydroxide) with water; the solution is called lime-water. Combines with carbon di-oxide, forming chalk (calcium carbonate)

Table of Common Oxides (*cont.*).

Metallic Elements	Oxide	Properties of Oxide Basic (see § 60). Counteract acids. If soluble, they produce *alkalies*
Magnesium	Magnesia	Very slightly soluble
Aluminium	Alumina	Insoluble. Combined with silica it forms many minerals and rocks
Iron	Red oxide of Iron	Insoluble. Combined with water it forms iron rust
Copper	Black oxide of Copper	Insoluble in water
Zinc	Zinc oxide	Insoluble; used as a pigment, Chinese white
Mercury	Red oxide of Mercury	Insoluble. Decomposes into mercury and oxygen on being heated
Lead	(1) Litharge (2) Red lead	Both insoluble in water (see § 37)

Questions on Chapter VIII.

1. What is the meaning of the word 'oxygen'? How is the gas usually prepared?

2. What do you mean by a 'catalytic agent'? Give an instance of the use of one.

3. If potassium chlorate contains 39·2 per cent. by weight of oxygen, how much of the salt must be taken in order to obtain 10 gm. of oxygen?

4. 8 gm. of potassium chlorate are heated till all the oxygen is driven off. What is the substance left? What will it weigh?

5. How would you prove that oxygen is more soluble than nitrogen in water?

6. What is an oxide? Describe the appearance of the oxides of magnesium, phosphorus, sulphur, calcium.

7. What is the effect on neutral litmus of the oxides produced by burning calcium, carbon, and iron in oxygen?

8. State the characteristics of metals. In what way do the metallic oxides differ from those of the non-metals?

9. (i) Calculate the weight of 1000 c.c. of oxygen S.T.P. (see § 46).

(ii) A litre of air (S.T.P.) weighs 1·293 gram. A litre of hydrogen (S.T.P.) weighs ·09 gram. Calculate the density of oxygen when (*a*) density of air = 1, (*b*) density of hydrogen = 1.

CHAPTER IX.

ACIDS—SULPHURIC, NITRIC AND HYDROCHLORIC.

52. We have learned that certain elements combine with oxygen; that many of the oxides produced are soluble in water; that solutions of the non-metallic oxides are acids. We have found too that Lavoisier was wrong in supposing that solutions of all oxides were acid and that the essential feature of an acid was that it should contain oxygen. About 35 years after Lavoisier's first work on oxygen, in 1810, Sir Humphry Davy showed that *not* all acids contain oxygen.

We have yet to learn that the essential feature of *acids* is that they all *contain an element called hydrogen* which may be *displaced* by a *metal* and that the formation of a class of substances called *salts* is the result of this displacement.

The three acids most commonly used in the laboratory were originally prepared from minerals and are called the 'mineral' acids to distinguish them from the 'organic' acids which were first known in plants and animals.

The Three 'Mineral Acids.'

Name	**Sulphuric Acid** or Hydrogen Sulphate or Oil of Vitriol	**Nitric Acid** or Hydrogen Nitrate or Aqua Fortis	**Hydrochloric Acid** or Hydrogen Chloride or Spirits of Salt
Density	1·8	1·5	1·2 (solution)
Properties	An oily heavy liquid, dissolves in water with production of great heat. Highly corrosive, b.p. 340° C.	A fuming brown liquid. Corrosive. Dissolves most metals. Stains skin yellow, b.p. 120° C.	A colourless gas, fuming in air and very soluble in water, (solution) b.p. 110° C.

53. Sulphuric Acid[1]. (For Preparation see Chap. XV.)

***Exp.** i. Dilute about 5 c.c. of concentrated acid by *pouring the acid in a thin stream into* about 20 c.c. of *water*, stirring all the time. [**Danger.** *Never* pour water on to strong sulphuric acid.] Record the rise in temperature.

***Exp.** ii. Make a thick syrup with a few lumps of sugar and a little water at the bottom of a large beaker. Pour a few c.c. of conc. sulphuric acid on the syrup and make a note of what you observe. The old-fashioned 'blacking' is made in this way, chalk being added to remove excess of acid.

***Exp.** iii. Write your name on a piece of paper, with the dilute sulphuric acid prepared in Exp. i, using a glass rod dipped in the solution. Gently evaporate off the water by warming the paper over the flame. Notice the charring of the paper and compare the result with Exp. ii, where *carbon* was also liberated.

***Exp.** iv. Pour *dilute* sulphuric acid on the metals zinc, iron and copper (in test tubes). Notice that zinc and iron dissolve and that an inflammable gas (*hydrogen*) is set free. Copper is not dissolved.

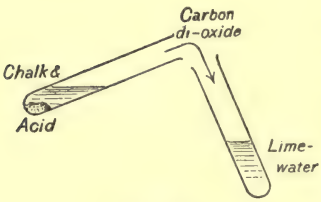


Fig. 34.

Exp. v. Warm, very carefully, *strong* sulphuric acid with the three metals named in Exp. iv, and notice that sulphur di-oxide (sulphurous anhydride) is liberated.

***Exp.** vi. Pour *dilute* sulphuric acid on small quantities of (*a*) washing soda, (*b*) chalk. Carbon di-oxide, a heavy gas, which extinguishes a lighted match and 'turns lime-water milky,' is set free. Pour the gas downwards into a tube containing clear lime-water and shake gently (Fig. 34).

54. Nitric Acid is prepared by distilling Chili saltpetre (sodium nitrate) or nitre (potassium nitrate) with strong sulphuric acid in iron retorts and condensing the fumes of nitric acid evolved.

***Exp.** i. Introduce by means of a paper funnel about 10 grams of nitre crystals to the bottom of a glass retort; set up the apparatus as shown

[1] **Danger.** Strong sulphuric acid should be used with *great care*.

in Fig. 35. Just cover the crystals with strong sulphuric acid [**Danger**], using a funnel to pour the acid through the tubulure of the retort. Warm the mixture gently with a small flame and condense the nitric acid fumes in a flask cooled by a running stream of cold water. Continue distilling until

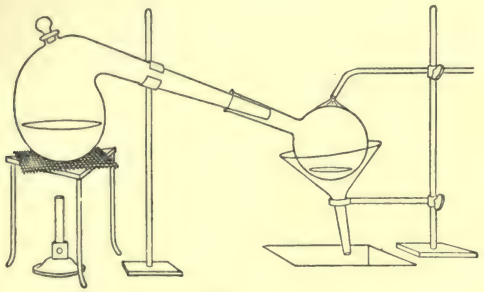


Fig. 35.

the fumes change from red to white. Remove the flask containing nitric acid, allow the retort to cool, then taking it by the stem carefully empty the liquid remaining in the retort into a porcelain evaporating dish. Notice that crystals are formed (potassium hydrogen sulphate). Wash these away with plenty of water.

Properties of Nitric Acid. Density about 1·45. Fumes in air. Highly corrosive. [**Danger.**] Destroys skin, flesh, clothes and other organic substances.

***Exp.** ii. Pour a little nitric acid on copper in a test tube. Notice that copper dissolves, that reddish fumes are set free, and that a blue solution is formed. Try other metals such as iron, zinc, tin and lead, and record what you observe in your note-book.

***Exp.** iii. Add nitric acid to (*a*) washing soda, (*b*) chalk, and notice that both dissolve, carbon di-oxide being liberated.

55. Hydrochloric Acid. (For preparation see Chap. XIV.)

***Exp.** i. Gently warm a few c.c. of conc. hydrochloric acid in a test tube. Notice that an acid gas is liberated from the solution.

***Exp.** ii. Pour the diluted acid on zinc, iron and copper. Compare the result of the experiment with Exp. iv, § 53.

***Exp.** iii. As in the case of sulphuric acid and of nitric acid, try the effect of adding hydrochloric acid to (*a*) washing soda, (*b*) chalk.

General properties of Acids.

(*a*) Diluted, acids have a sour taste, and change the colour of certain '*indicators*' (*e.g.* blue litmus to red).

(*b*) Corrode or dissolve certain metals with liberation of a gas or gases (usually hydrogen).

(*c*) Dissolve washing soda and chalk, liberating carbon dioxide.

Later we shall discuss more fully the statement made at the beginning of this chapter—that *all acids contain hydrogen* and that when a metal dissolves in acids the hydrogen of the acid is turned out and a *salt* is formed (cf. Exp. ii, § 31).

Questions on Chapter IX.

1. What tests would you apply to discover whether a given acid was sulphuric, nitric or hydrochloric?

2. When mixing strong sulphuric acid and water why is the acid always poured slowly into the water instead of water into acid?

3. What difference is there in the action of strong and dilute sulphuric acid on copper?

4. What is the effect of adding acids to chalk? How would you identify the gas evolved?

5. How may nitric acid be prepared?

CHAPTER X.

BASES (INCLUDING ALKALIES) AND SALTS. WATER OF CRYSTALLIZATION.

56. Alkalies are the soluble members of a larger class called **Bases.** When we burned the metals sodium, potassium, calcium and magnesium in oxygen, we found that the oxides formed were more or less soluble in water. The solutions of metallic oxides are soft or soapy to the touch, turn red litmus blue and destroy acids. They are called **alkalies** and are said to be *alkaline* in reaction. Because they are combined with *water* (Greek 'hudor') they are named metallic *hydrates* or *hydroxides.*

The **Caustic alkalies.** *Sodium hydrate* and *potassium hydrate* are *caustic, i.e.* they *burn* or corrode the skin and flesh, consequently they are frequently called caustic soda and caustic potash respectively.

Calcium hydrate (slaked lime) and its solution (lime water) are not caustic.

Another common alkali, **ammonium hydrate,** is a solution of the gas *ammonia* in water.

57. Preparation of Ammonia.

*Exp. i. Mix about 10 grams of *sal ammoniac* (ammonium chloride) crystals with about the same weight of powdered *quicklime* (calcium oxide) and place the mixture in a wide tube; add a few grams of quicklime and fit the tube, sloping slightly downwards, with a cork and upward delivery tube

(Fig. 36). Warm the mixture gently, keeping the flame moving, and collect two cylinders of the gas by 'upward displacement.'

(1) Plunge one cylinder mouth downwards into a basin of water. The water rushes up and nearly fills the jar. At 0° C. one volume of water dissolves 1100 volumes of ammonia.

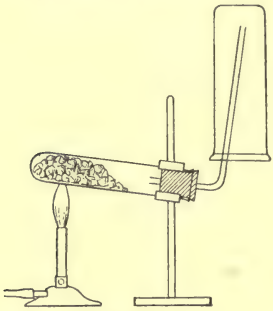


Fig. 36.

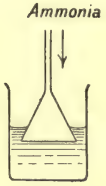


Fig. 37.

(2) Pour a few drops of strong hydrochloric acid into a gas jar and fit the second cylinder of ammonia to it, mouth to mouth. Dense fumes of solid ammonium chloride are formed.

(3) Turn the delivery tube so that it points downwards and fit a funnel to it. Pass ammonia into water, keeping the funnel's mouth just below the surface of the water (Fig. 37). The solution obtained is strongly *alkaline* and is lighter than water. The specific gravity of a saturated solution of ammonia is ·880.

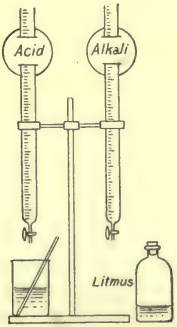


Fig. 38.

***Exp.** ii. Show that ammonia is liberated when sal ammoniac (or any ammonium compound) is heated with caustic soda, potash or lime.

58. Neutralizing Acids with Alkalies. We have already found that a definite quantity of acid at a certain strength will always dissolve the same weight of metal (§ 31, Exp. ii). It is probable therefore that *acids and alkalies of fixed strength will 'destroy' each other in definite proportions.* Let us try whether this is so.

***Exp.** Fill two burettes (Fig. 38), the one with

dilute *caustic soda* solution (sodium hydrate), the other with dilute *hydrochloric acid* (spirits of salt). Adjust the levels to zero in your burettes. Run about 5 c.c. of alkali into a beaker, add 2 or 3 c.c. of neutral *litmus*, then add acid cautiously until the solution just turns reddish, stirring all the time with a glass rod. Record the readings. (Avoid *parallax* error.) Add about 5 c.c. of alkali at a time and continue the process of *neutralizing* with acid, *i.e.* running in acid or alkali as is needed to obtain a solution which is *neither* blue nor red. Record and work out your results as follows:

	Burette Readings		Proportion $\frac{\text{Alkali}}{\text{Acid}}$	Vol. Alkali neutralizing 1 c.c. of Acid	
	Alkali	Acid			
1	5 c.c.	4 c.c.	$\frac{5}{4}$	1·25 c.c.	The average of these results shows that 1·25 c.c. (nearly) of this particular solution of caustic soda will neutralize 1 c.c. of this particular solution of hydrochloric acid
2	10·1 ,,	8·1 ,,	$\frac{10·1}{8·1}$	1·25 ,,	
3	14·9 ,,	11·8 ,,	$\frac{14·9}{11·8}$	1·26 ,,	
4	20·5 ,,	16·5 ,,	$\frac{20·5}{16·5}$	1·24 ,,	
5	26 ,,	21 ,,	$\frac{26}{21}$	1·24 ,,	

59. Preparation of a neutral Salt. Multiply the results in the fourth column above by 10. Then 12·5 c.c. of our caustic soda solution added to 10 c.c. of hydrochloric acid solution will give us a neutral solution and we need not add litmus to prove it.

***Exp.** i. From the average of your results calculate how much of your two solutions you must add together to produce about 20 c.c. of neutral solution. Run the quantities into an evaporating dish. Test whether the solution is neutral by putting a drop on red and blue litmus papers; then evaporate the solution to dryness. The white solid obtained is easily recognized as common **salt**, *sodium chloride*.

***Exp.** ii. Repeat Exp., § 58, and the last experiment, but use dilute solutions of *caustic potash* (potassium hydrate) and *nitric acid*. Evaporate

nearly to dryness. Needle-shaped crystals of the **salt,** *potassium nitrate,* will form in the neutral solution, on cooling.

***Exp.** iii. Neutralize ammonia solution with hydrochloric acid; evaporate until crystals form on cooling. Dry the crystals of the salt, *ammonium chloride,* between filter paper. Cf. § 57 (2).

If these dry crystals are heated in a test tube, they *sublime, i.e.* they become gaseous and reform on the cool sides of the tube. The process is called *sublimation.*

60. Metallic oxides, insoluble in water, may dissolve in acids and form salts. We have learned that the solutions of metallic oxides (alkalies) neutralize acids to form salts. Let us try whether metallic oxides, such as black copper oxide and litharge (oxide of lead), which are insoluble in water, will dissolve in acids, and whether a salt is formed in solution.

***Exp.** i. Pour into an evaporating dish about 20 c.c. of dilute *sulphuric acid* and stir in *black oxide of copper,* warming the mixture and gradually adding the oxide until a black residue remains undissolved. The copper oxide is then said to be '*in excess,*' *i.e.* there is more than is needed to complete the reaction. Filter the solution and evaporate until crystals form. Notice that the crystals are greenish blue. They are called **copper sulphate.** Dry them between filter paper and keep them for future use.

***Exp.** ii. In exactly the same way prepare crystals of **lead nitrate** by adding litharge (yellow oxide of lead) to dilute nitric acid.

In each of these two experiments test with litmus paper whether the solution from which crystals of the salt form is neutral. In § 59 we prepared neutral salts but we shall find that some salts are alkaline and others acid in reaction.

***Exp.** iii. Heat *strongly* dry crystals of copper sulphate in a hard glass tube. Black oxide of copper is recovered and sulphurous fumes are given off.

***Exp.** iv. Heat dry crystals of *lead nitrate.* Litharge is recovered and nitrous fumes as well as oxygen are liberated.

Bases. The earthy powder that is obtained by heating a salt to a high temperature was known by the word *base* or *foundation* of the salt. We now apply the term *base* to a metallic oxide, soluble or insoluble in water. The *soluble* bases we call **alkalies** and these are the only bases which change the colour of litmus and other indicators.

Bases, therefore, are *metallic oxides*, whether soluble or insoluble in water, which combine with *acids* to form *salts*.

Copper oxide then dissolved in acids will produce *copper salts*; calcium oxide, *calcium salts*; lead oxide, *lead salts*.

But various bases dissolved in:

Sulphuric acid	will yield	salts	called	metallic *sulphates*,
Nitric acid	„	„	„	*nitrates*,
Hydrochloric acid	„	„	„	*chlorides*.

61. Water of Crystallization. When we heated copper sulphate crystals (§ 60, Exp. iii) we probably noticed that steam was given off. Water enters into the composition of many crystalline salts. If the salt is gently heated the crystal splits up into water and the **anhydrous** (= without water) salt.

***Exp.** i. Gently heat the following crystalline salts in dry test tubes, previously grinding the salts to a fine powder in the mortar: (*a*) copper sulphate, (*b*) nickel sulphate, (*c*) washing soda.

***Exp.** ii. **To find the percentage of water of crystallization in crystals of Barium Chloride.** Grind the crystals to a fine powder. Weigh a crucible and into it weigh about one gram of the powdered Barium Chloride. Heat this gently (Fig. 39) for about ten minutes, allow it to cool and weigh again. Repeat this process until two consecutive weighings are the same. Note the loss in weight and calculate the weight of water which 100 grams of the crystals would lose. A good result gives water of crystallization 14·8 %.

Fig. 39.

If time permits prove by experiment that copper sulphate contains about 36 % of water of crystallization.

Questions on Chapter X.

1. What do you mean by a 'base' and an 'alkali'? What are the characteristics of alkalies?

2. How would you show that ammonia gas (*a*) is lighter than air, (*b*) is very soluble in water?

3. Describe how to obtain all the alkalies mentioned in § 56.

4. How could you prepare a specimen of ordinary salt from solutions of caustic soda and hydrochloric acid?

5. Describe the substances obtained when (*a*) dilute sulphuric acid acts upon black oxide of copper, (*b*) nitric acid acts upon litharge. What is the effect of strongly heating the two products?

6. What takes place when a solution of potash is neutralized by adding nitric acid? Describe one or two simple experiments to demonstrate the nature of the body formed. (O. J.)

7. Describe, with sketch of apparatus, how a jar of *dry* ammonia gas may be obtained.

8. What is meant by 'water of crystallization'? Find the percentage of water of crystallization in the following substances:

	(A) Barium Chloride	(B) Copper Sulphate
Weight of crucible empty	22 gm.	21·8 gm.
,, ,, + salt before heating	22·5 ,,	25·6 ,,
,, ,, + residue after heating	22·425 gm.	24·2 ,,

CHAPTER XI.

HYDROGEN AND WATER.

62. The Hon. Henry Cavendish in 1766 published a paper describing how he had prepared 'Inflammable Air' (Hydrogen) by the action of acids on metals. We shall find later that Cavendish was wrong when he conjectured that the gas came from the metal. Already we have found that all the three acids with which we are acquainted dissolve metals and that in the majority of instances an 'inflammable air' is given off. It is probable that the gas comes from the acid, as the same acid acting on different metals yields the same gas (§§ 53 and 55).

***Exp.** i. **Preparation of Hydrogen.** Use the apparatus shown in Fig. 40. The cork has two holes, the one for the delivery tube, the other for a funnel tube down which *dilute sulphuric acid* may be poured on to *granulated zinc*. Collect several cylinders full of hydrogen by water displacement.

Fig. 40.

***Exp.** ii. Hold a lighted taper vertically and bring the first cylinder mouth downwards over the flame. The slight explosion shows that *hydrogen mixed with air is explosive*. Repeat the experiment with the second and the third jars. With the third the explosion will not be so violent. Hydrogen burns with a *light blue flame*. Push the taper into the jar; the light is extinguished.

***Exp.** iii. *Hydrogen is the lightest substance known.* Try whether you can pour the gas from the fourth cylinder into another jar held mouth downwards (Fig. 41). Again test with the lighted taper.

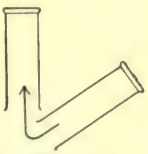


Fig. 41.

Exp. iv. Using a second thistle funnel fitted to the end of the delivery tube, blow *soap bubbles* with hydrogen. Obtain the necessary pressure by closing the funnel, which passes through the cork, with the hand or with a cork. If the bubbles do not form readily, remove all traces of acid from the gas by passing it through a solution of soda.

Hydrogen is used to fill balloons and dirigibles.

A litre (1000 c.c.) of hydrogen S.T.P. weighs ·09 gram.

63. Water is formed when Hydrogen burns in air.

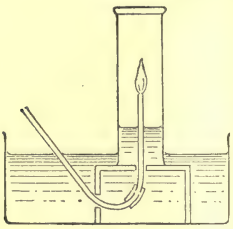


Fig. 42.

***Exp.** i. Collect a test tube of the gas by upward displacement and try whether it burns quietly. If there is no explosion, it will be safe to put a light to the end of the delivery tube. Lengthen this tube until its end stands about 4 inches above the surface of the water in the trough (Fig. 42). Place a dry cylinder full of air over the flame of hydrogen. The water rises in the cylinder and the flame goes out (cf. § 15). Test the residual gas with a lighted taper. Evidently oxygen has been removed by combination with hydrogen. On the sides of the jar is seen a deposit of moisture. Let us proceed to collect the liquid which has been formed.

***Exp.** ii. First remove water vapour from the hydrogen by inserting a drying tube containing *calcium chloride*[1] or pumice stone saturated with strong *sulphuric acid*[1]. Light the *dry* hydrogen issuing from the delivery tube and let the flame impinge against the side of a flask, kept cool by a stream of water as in Fig. 43. Very slowly a liquid condenses. Bear in

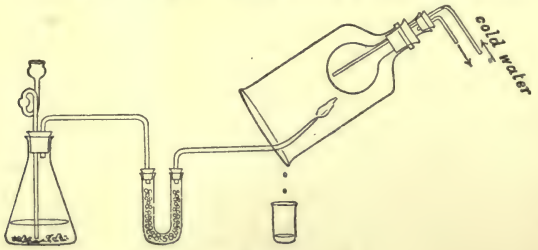


Fig. 43.

[1] Both substances absorb moisture.

mind the lightness of hydrogen; you will have to burn 11 litres before a gram of the gas has been used up. In 15 minutes a dozen such pieces of apparatus will yield together sufficient *liquid* to test that it freezes at 0° C., boils at 100° C., that its density is unity, that it is tasteless and colourless and that it leaves nothing on evaporation—in short *the liquid is water.*

64. Hydrogen combines with oxygen in some metallic oxides. Reduction. (See also Chapter XVI.)

***Exp.** Use the same apparatus as before for obtaining a current of dry hydrogen. Pass the gas over heated (1) yellow oxide of lead (litharge), (2) red oxide of lead (red lead), (3) oxide of iron, (4) oxide of copper and condense the *water* formed in each case in a U tube as shown in Fig. 44.

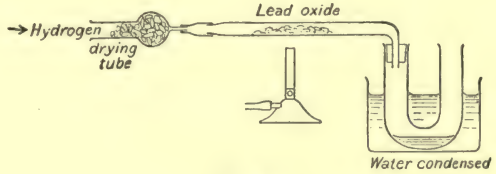


Fig. 44.

In each case the metallic oxide is turned or 'led back' again (*reduced*) to metal. The *lead* may be recognized by its malleability, its softness (it marks paper) and its readiness to melt and tarnish on being heated in air. The powdery *iron* is attracted by a magnet and dissolves in sulphuric acid with liberation of hydrogen. We recognize the *copper* by its reddish golden colour and by the fact that it dissolves in nitric acid producing a blue solution. The *condensed liquid* in the U tube answers all the tests for *water* given in § 63, Exp. ii.

We may therefore conclude (1) that *water* consists of *hydrogen* and *oxygen*, (2) that hydrogen can *reduce* certain metallic oxides to their metals, water being formed at the same time.

Other metals such as calcium, magnesium and sodium, combine so strongly with oxygen that hydrogen will not reduce their oxides. No effect is noticed if lime, for instance, is heated in a stream of hydrogen. Calcium and oxygen have such 'affinity' that the presence of hydrogen does not disturb their combination.

65. Composition of Water by Weight.

***Exp.** Again use the same flask and drying tube for preparing hydrogen but insert a second drying tube to remove any traces of moisture. Black oxide of copper carefully dried is placed in a porcelain 'boat' and the weight obtained [=(say) 8·12 grams]. The 'boat' and oxide *A* are then placed in a piece of hard glass tubing drawn out so as to fit into a small flask or bulb (*B*) connected with a calcium chloride tube (see Fig. 45) which prevents loss of water vapour.

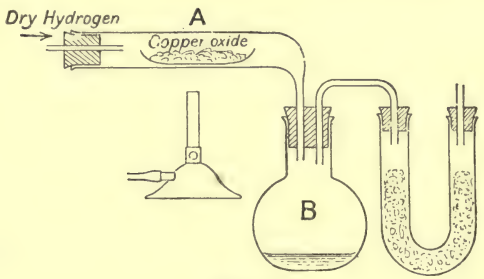


Fig. 45.

The small flask, cork and calcium chloride tube *B* are weighed *together* [=(say) 49·52 grams). The hard-glass tube is now fitted to the small flask and dry hydrogen is then passed *slowly* through the apparatus. Begin heating the copper oxide near the bend so that all moisture is carried over into *B*. After reduction seems complete, stop heating but continue to pass hydrogen until the copper is cold. Then detach the hydrogen generating flask and sweep out the hydrogen from *A* and *B* by means of a slow current of dry air from an aspirator. Weigh the 'boat' and reduced copper *A* [=(say) 6·92 grams]. Weigh the flask and U tube *B* again [=(say) 50·87 grams].

The weight of water formed = gain in weight of $B = 50{\cdot}87 - 49{\cdot}52 = 1{\cdot}35$ grams of water which contains the weight of oxygen lost by $A = 8{\cdot}12 - 6{\cdot}92 = 1{\cdot}20$ grams of oxygen,

$\therefore 1{\cdot}35 - 1{\cdot}20 = 0{\cdot}15$ gram = the weight of hydrogen,

$$\therefore \frac{\text{weight of oxygen}}{\text{weight of hydrogen}} = \frac{1{\cdot}20}{0{\cdot}15} = \frac{8}{1}.$$

66. Hydrogen from Water and Steam. The metals potassium, sodium, and calcium, have such an 'affinity' for, or eagerness to combine with, oxygen, that they liberate hydrogen

when placed in contact with water. The metallic oxide that is formed combines with more water to form the metallic hydroxide or hydrate. We have found, in these three instances (§ 56), that the hydroxide is soluble in water and therefore its solution gives an alkaline solution.

Exp. i. Drop small pieces of **potassium** (about the size of half a pea) on water in a dish. [**Danger**, see § 23.] Notice that the hydrogen gas liberated burns with a violet flame and that a soapy, alkaline liquid is formed (test with litmus).

Exp. ii. Repeat, substituting **sodium** for potassium. The action is not so violent but still there is **Danger**. The hydrogen liberated will catch fire if the sodium is dropped on to filter paper floating on the water[1].

Exp. iii. Drop a few turnings of **calcium** into water in a dish and place over them a cylinder filled with water (Fig. 46). Hydrogen collects and may be tested with a burning taper. Calcium oxide (quicklime) is probably first formed, but this combines immediately with water to form calcium hydroxide (slaked lime) which is slightly soluble in water, hence the solution is alkaline.

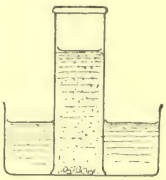


Fig. 46.

***Exp.** iv. **Magnesium** will not combine with the oxygen in water unless the water is presented to the heated metal in the form of steam. Blow a small hole in a hard glass test tube, having softened the end of the tube in a blow-pipe flame. Put *magnesium* turnings in the tube and pass a current of steam through it. Place a sand tray under the tube (Fig. 48). Now heat the magnesium with the Bunsen flame. Hydrogen is liberated and catches fire as it issues from the tube. At the same time magnesium oxide is formed.

[1] The safest and most striking method of showing the action of sodium on water is as follows. Clamp a piece of $\frac{5}{8}$" tubing vertically with its end dipping into water containing a little of the 'indicator,' phenolphthalein (Fig. 47). Roll a piece of clean sodium, about half the size of a pea, between the finger and thumb (*dry*) and drop the little ball of sodium obtained down the tube. Hydrogen is liberated and may be burned at the mouth of the tube and at the same time the presence of sodium hydroxide is shown by the solution turning pink.

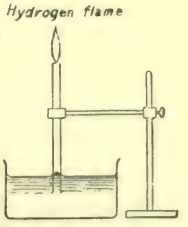


Fig. 47.

We remember that magnesium oxide is very slightly soluble in water and produces a very weak alkaline solution. (Calcium and zinc may be substituted for magnesium in this experiment.) (Fig. 48.)

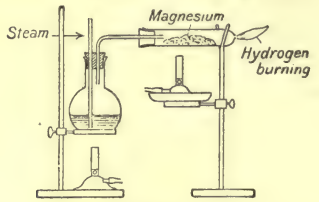


Fig. 48.

A reversible change. *Iron heated to redness combines with the oxygen in steam* provided that the hydrogen set free is swept away by the current of steam. Otherwise the reaction tries to *reverse* itself, for we have seen already that hydrogen *reduces* oxide of iron back to iron with the formation of water at the same time.

Exp. v. Pass a current of steam through a red-hot iron tube containing iron nails and collect the liberated hydrogen in a pneumatic trough. (Fig. 49.) When sufficient hydrogen has been collected, reverse the process

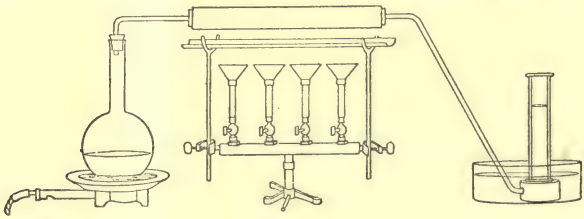


Fig. 49.

by passing dry hydrogen through the heated tube and condense the steam that is formed in a cooled U tube (Fig. 50).

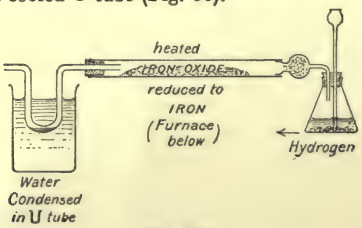


Fig. 50.

This reversible change may be indicated in the following way—

Oxide of hydrogen (steam) + iron $\rightleftarrows$ oxide of iron + hydrogen.

If a mixture of steam and hydrogen is passed through the red-hot iron tube both iron and oxide of iron are found in the tube. If the quantity of steam is increased more oxide of iron is formed, but if the quantity of hydrogen is increased more oxide of iron is reduced to iron. The *direction* of the reaction therefore depends on the *weight* of the reacting substance present.

67. Electrolysis of Water *containing sulphuric acid.* The process of splitting up or dividing a body into its constituent parts is called **analysis**. If this analysis is effected by means of an electric current the method is called **electrolysis**. Pure water is not a good conductor of electricity, but if a few drops of an acid or a solution of a metallic salt be added the current passes and at the same time the substance in solution may be split up. The reason for this is not very simple but the effect is easily brought about.

Exp. Two plates of **platinum** called *electrodes* are each connected with wires which pass through a broad rubber stopper fitted to a small trough (Fig. 51) containing *water* and a few c.c. of *dilute sulphuric acid*. A tube filled with water is inverted over each platinum electrode, and the ends of the wires are connected with the poles of a battery of 3 or 4 bichromate cells or accumulators. When the circuit is completed, gases are liberated at the electrodes in the proportion by volume of 2 to 1. On testing the gases with a lighted match, the *two volumes* are found to be *hydrogen*, and the *one volume, oxygen*.

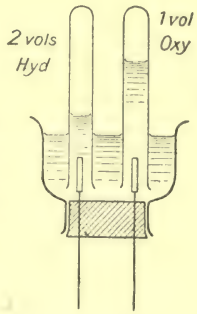


Fig. 51.

Now if the density of hydrogen = 1 and the density of oxygen = 16, the two volumes of hydrogen will weigh 2 and the one volume of oxygen will weigh 16. That is the weight of the hydrogen liberated is to the weight of the oxygen liberated in the proportion 1 to 8 (cf. § 65).

68. Synthesis of Water. Synthesis means a 'placing together.' If hydrogen and oxygen are 'placed together' in a strong vessel in the *proportions of* 2 *to* 1 *by volume* no change

occurs until a spark or flame is brought into the mixture. Then a violent explosion takes place and the two gases combine with great heat to form water. The Hon. Henry Cavendish, the discoverer of 'inflammable air,' was the first to perform this experiment (about 1785) and to find the exact proportions in which the two gases combine. He made use of his discovery to determine the amount of gas available for combustion or for breathing in the atmosphere. If air is burned or exploded with sufficient hydrogen to combine with all the oxygen present, one-third of the volume of the two gases combining must be oxygen. The total volume of hydrogen and oxygen taking part in the explosions is easily measured by the contraction which occurs. Cavendish called the instrument in which he estimated the quality of the air a 'goodness-measurer' or **eudiometer.** The most convenient form is shown in Fig. 52. A long U tube of thick glass, one limb of which is closed and through which two platinum wires are fused, is filled with mercury and the gases to be exploded are introduced by a long delivery tube passed down the open limb. The level is adjusted by pouring mercury in at the top of the open limb or by running it out at the tap as required. The explosion is effected by sending an electric spark across the gap between the two wires, and loss of mercury is prevented by closing the open end of the U tube with the thumb.

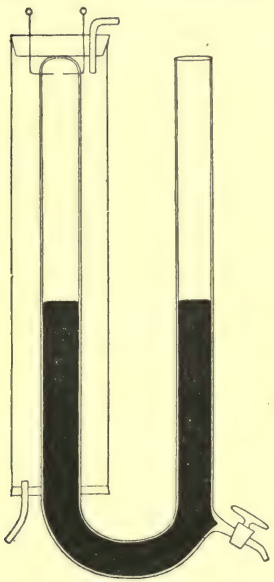


Fig. 52.

If the closed limb is surrounded by a jacket through which the vapour of a liquid boiling above 100° C. is passed, steam

(gaseous water) is formed after explosion. If the conditions of temperature and pressure are unaltered during the experiment, it is found that *two volumes of hydrogen combine with one volume of oxygen to form two volumes of steam.*

Exp. Measure the volume of a *soda-water bottle* by filling it with water from a measuring cylinder. Then mark with a file, on the bottle, *two-thirds* of the volume. Fill the bottle with water, and, using the pneumatic trough, pass *hydrogen* until the bottle contains $\frac{2}{3}$ of its volume of the gas. Displace the remaining $\frac{1}{3}$ of water with *oxygen* and cork up the mixed gases. Wrap the bottle in a damp cloth as a precaution against breakage and uncork the bottle at a flame. A loud explosion is heard.

Repeat, using $\frac{2}{7}$ of air and $\frac{5}{7}$ of hydrogen, and compare the violence of the report.

69. Sufficient experiments have now been made to justify us still further in concluding that:

(*a*) *Water is a compound of hydrogen and oxygen.* [See § 24.]

(*b*) *Hydrogen is contained in the acid* and not in the metal. [See §§ 62 and 67.]

(*c*) The name *hydrogen* or '*water producer*' is not wrongly chosen

70. Water is formed when a basic oxide neutralizes an acid. In studying the interaction of *acids* and *bases* we have confined our attention to the formation of *salts*. We have learned that the *same* salt is produced when an acid is neutralized

(*a*) by dissolving a metal in the acid,

(*b*) „ „ metallic *oxide* in the acid;

the only difference being that in the first case hydrogen is generally liberated but not in the second. Having now established the fact that *water* is an *oxide* of *hydrogen*, we are probably right in surmising that the *hydrogen of the acid* and the *oxygen of the base* combine to *form water* in the second case. The proof of this hypothesis however is reserved until, having prepared a *dry*

current of an acid gas, we can pass it over a *dry* metallic oxide and show that *water* is formed as well as a salt. [§ 93 (4).]

Questions on Chapter XI.

1. Describe how hydrogen is usually prepared, giving a sketch of apparatus. By what tests would you identify the gas?

2. How would you prove that when hydrogen is passed over heated iron oxide or lead oxide metallic iron or lead is obtained and water formed?

3. Describe fully how the composition of water by weight may be found.

4. What do you mean by a 'reversible change'? Give an example.

5. Describe an experiment by which water may be synthesised.

6. What differences are there in the action of sodium, potassium, and calcium on water?

7. State what is meant by 'electrolysis' and describe in detail the electrolysis of water.

8. If 15 c.c. of hydrogen and 10 c.c. of oxygen are mixed and exploded in a eudiometer how much gas will be left? Of what will it consist?

9. If 9 gm. of water are decomposed what weight and volume of hydrogen will be obtained at S.T.P.?

10. How would you prove that there is a contraction in volume when hydrogen is exploded with or burnt in air? (C. J.)

11. Name the metals mentioned in § 66 in the order which shows their relative affinity for (or eagerness to combine with) oxygen. Give reasons.

CHAPTER XII.

CHALK, LIME, AND CARBONIC ACID GAS, NATURAL WATERS.

71. The first part of the title of this chapter written in chemical terms is—*Calcium Carbonate*, *Calcium Oxide*, and *Carbon dioxide*. Already we are acquainted with these substances.

Carbon dioxide, or *carbonic anhydride*, or *carbonic acid gas* was first known to us when we burned carbon (charcoal) in oxygen gas. It was found to be a colourless gas, which extinguishes a taper and 'turns lime water milky.' Shaken with water, it dissolved and formed a very weak acid solution. Consequently we should expect it to combine with alkalies such as potash, soda and lime to form salts. These *salts* are called **Carbonates.**

Calcium Oxide (*quicklime*) was formed when we burned the metal *calcium* in oxygen. It was found to be a white solid which combines with water to form a slightly soluble *hydrate* or hydroxide (slaked lime) and the solution—*lime water*—we know to be a weak alkali. Calcium oxide, a metallic oxide, is a *base* and will neutralize acids to form salts, as we have seen above, where, with the weak carbonic acid it yielded calcium carbonate.

Calcium Carbonate (*carbonate of lime*) occurs in nature in great abundance and in many varieties. Chalk, limestone, marble, calc-spar, sea-shells, coral, egg-shells, all are more or less pure calcium carbonate. The chalk cliffs and hills of the South and East Coasts and of the Eastern Midlands of England, and the limestone of the Pennines are masses of rock, in some places several thousands of feet in thickness, formed by the aid of animals and plants which lived in the oceans countless ages ago, before the present continents existed. The process of formation of limestone is going on to-day in many parts of the sea.

Now as we have some idea of the chemical composition of calcium carbonate, a little thought will suggest methods of dividing it into its two components—calcium oxide (lime) and carbon dioxide (carbonic acid gas). The latter substance is so weak an acid-forming oxide that we should expect that the two oxides are not very firmly combined—

(*a*) Will heat cause the basic oxide and acid oxide to part company?

(*b*) Will a strong mineral acid combine with the base and liberate the acid oxide?

72. Action of Heat on Calcium Carbonate. Lime is used in large quantities for making cement and mortar. In many parts of the country 'lime burning' is carried on in pit-shaped furnaces called **lime-kilns** (Fig. 53) where limestone (calcium carbonate) is mixed with coal and heated until the mass is red-hot. The limestone loses nearly half its weight and a white substance called quicklime (calcium oxide) remains. Quicklime combines with nearly one-third of its weight of water in being converted to **slaked lime.**

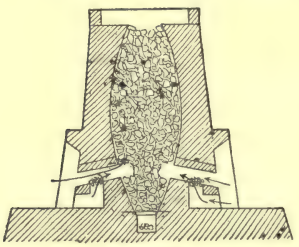


Fig. 53.

***Exp.** i. Grind about 2 grams of powdered calc-spar to a fine dust in an agate mortar. Weigh a small crucible and into it weigh *one gram* of the *calcium carbonate*. Heat the crucible and calc-spar for half an hour over the blowpipe flame or in a muffle furnace. Allow the crucible to cool in the desiccator and weigh again. 0·56 *grams of quicklime* remain in the crucible, *i.e.* there is a loss of weight = 0·44 gram.

Exp. ii. While Exp. i is proceeding, heat strongly about ½ a gram of calc-spar in a hard glass tube over the blowpipe flame and draw any gas liberated through a wash bottle containing lime water (Fig. 54), with the aid of an aspirator. A gas is liberated which forms a white precipitate with the lime water.

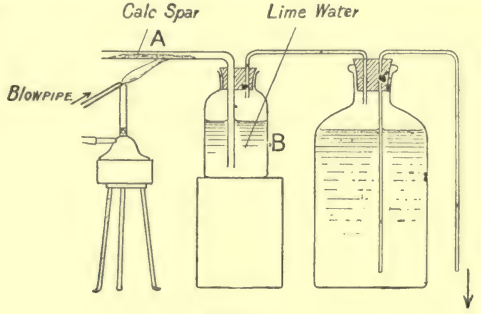


Fig. 54.

If a large class of students is performing this experiment a sufficient quantity of the precipitate may be collected, dried and heated similarly in this apparatus (Fig. 54). A similar result is obtained proving that the precipitate formed is calcium carbonate.

In *A* the carbonate has been split up (*analysed*); in *B* the carbonate has been formed again by the placing together (*synthesis*) of its components. We have therefore found that calcium carbonate (*a*) yields quicklime and carbonic acid gas on heating (analysis), (*b*) is formed by the union of lime and carbonic acid gas (synthesis).

Exp. iii. Pour water on a lump of *quicklime* (calcium oxide) contained in a beaker. In a few minutes much heat is evolved and a powdery substance, *slaked lime* (calcium hydrate or hydroxide), is formed. Shake the slaked lime with water in a large bottle and allow it to stand. The clear solution is called *lime water* and has an *alkaline* reaction.

73. Action of Acids on Calcium Carbonate. The fact that calcium carbonate is composed of a base and a weak acid oxide suggests that probably one of the stronger acids will decompose it by combining with the base to form a salt, at the same time the gas (carbon dioxide) will be liberated.

Chalk is a typical carbonate and its action towards acids is also typical. The metallic base which it contains will neutralize an acid and form a salt just as a simple base does, but in addition the gas (carbon dioxide) is set free.

We have already found that mineral acids cause an effervescence of gas when poured on the carbonates of soda, potash and lime (§ 53). Let us now proceed to find whether the various forms of calcium carbonate lose the same percentage by weight (44 %, see § 72, Exp. i) when an acid is added as they did when strongly heated.

74. To find the percentage loss of weight and to measure the volume of gas set free when acid is added to calcium carbonate. Exactly *one gram* of powdered *calc-spar* is weighed 'by difference' into a small short test tube. Carefully lower the test tube into a small light flask previously fitted up as in Fig. 55. The flask contains about 20 c.c. of concentrated *hydrochloric acid* and an equal volume of water. The right-hand tube reaches nearly to the level of the acid and is closed at the top by a stopper. The larger tube is packed with cotton wool or calcium chloride. The apparatus shown in Fig. 55 is carefully weighed and then fitted to the aspirator[1]. The flask is now tilted slightly until the acid dissolves the

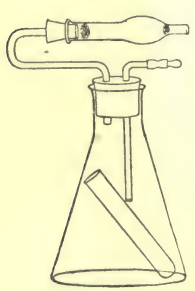


Fig. 55.

[1] A layer of paraffin oil on the surface of the water in the aspirator prevents loss of volume due to the slight solubility of carbon dioxide.

calc-spar. When the volume of gas set free has been *measured* by the *quantity of water displaced from the aspirator*, the calcite being completely dissolved, the stopper of glass rod is removed from the top of the narrow tube and air is drawn through the flask until the latter is again full of air. The stopper is then replaced and the flask and its contents are weighed again. Accurate experiment shows that the *loss of weight* = 0·44 *gram*, a result corresponding to the loss of weight when a gram of calc-spar was strongly heated (§ 72, Exp. i).

Other forms of calcium carbonate give the *same percentage composition*. We conclude then that calc-spar, calcite, marble, and chalk are all variations of *calcium carbonate* and that this substance is a *chemical compound* consisting of 56 °/₀ by weight of *calcium oxide* and 44 °/₀ of *carbon dioxide*.

In an experiment the *volume* of the gas displaced from 1 gram of calc-spar was 240 c.c. The *temperature* was 15° C., the *height of barometer* was 752 mm. Pressure of aqueous vapour at 15° C. = 12 mm.

∴ at S.T.P. the volume of carbon dioxide (see § 9)

$$= \frac{240 \times 273 \times (752 - 12)}{(273 + 15) \times 760} = 222 \text{ c.c.}$$

∴ 222 c.c. of carbon dioxide (S.T.P.) weigh 0·44 gram.

∴ **1000** c.c. „ „ „ $\frac{0{\cdot}44 \times 1000}{222}$ gram

$= \mathbf{1{\cdot}98}$ gram.

Density of a gas

$$= \frac{\text{Wt. of a given volume of gas}}{\text{Wt. of same volume of hydrogen (under same conditions)}}.$$

∴ Density of carbon dioxide

$$= \frac{\text{Wt. of 1000 c.c. carbon dioxide S.T.P.}}{\text{Wt. of 1000 c.c. hydrogen S.T.P}}$$

$$= \frac{1{\cdot}98 \text{ gram}}{0{\cdot}09 \text{ gram}} = \mathbf{22}.$$

75. Preparation of Carbonic Acid Gas and Calcium Chloride.

***Exp.** i. Fit up the same apparatus as was used in the preparation of hydrogen. Small lumps of *marble* are placed in the flask and are covered with water. Pour a little strong *hydrochloric acid* down the funnel. Collect four cylinders full of the carbon dioxide set free; then pass the gas through about 60 c.c. of *lime water* in a boiling tube. Notice that a white precipitate (calcium carbonate) is formed (cf. § 72, Exp. ii). Allow the gas to bubble through the milky liquid for several minutes.

***Exp.** ii. Plunge a lighted taper into one of the jars of carbon dioxide—the taper is extinguished.

***Exp.** iii. Pour the carbon dioxide gas downwards from one cylinder into another containing air. Test the gas in the second cylinder with a burning taper or with lime water. This experiment suggests that we may collect the gas not by *water* displacement but by *air* displacement (see Fig. 56).

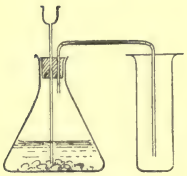


Fig. 56.

***Exp.** iv. Plunge a piece of burning *magnesium* ribbon into a jar of the gas[1]. Notice that magnesium oxide is formed and that black specks of *carbon* are liberated. Here carbon dioxide has been decomposed (analysis) into free carbon and oxygen, the latter having combined with magnesium. Contrast this experiment with the *synthesis* (putting together) of the gas when charcoal was burned in oxygen (§ 49).

***Exp.** v. Pour *caustic soda* or *caustic potash* solution into the fourth cylinder of carbon dioxide and shake. Notice the gas is absorbed forming the *carbonate* of soda or potash according to the alkali used. Add a few drops of acid to the solution and notice that carbon dioxide is set free.

76. Temporarily Hard Water. The carbonic acid gas has now been bubbling through the lime water (§ 75, Exp. i) for some minutes and probably the precipitate of calcium carbonate has dissolved. The clear solution is called Temporarily Hard Water and contains a substance, **calcium bi-carbonate,**

[1] This result is better shown by passing dry carbon dioxide over magnesium powder carefully heated in a bulb tube (§ 80, Fig. 58). A larger quantity of carbon may be thus obtained.

which will not lather with **soap.** Divide the 60 c.c. of hard water into *three equal parts* in separate small flasks *a*, *b* and *c*.

***Exp.** i. (*a*) To the *first* add *lime water* slowly, shaking constantly, *until* there is a thick white precipitate formed. *Do not add excess of lime water.* (*b*) Boil the *second* quantity of hard water and show that carbon dioxide gas is driven off by passing the gas into lime water (Fig. 57). Notice that the calcium carbonate appears as a white precipitate when the gas has been driven off by boiling.

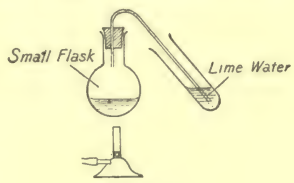


Fig. 57.

***Exp.** ii. Set up a burette filled with a solution of *Castille soap* in methylated spirit and run the soap solution into the three flasks containing respectively (*a*) hard water to which lime water has been added; (*b*) hard water, boiled; (*c*) the temporarily hard water, untouched. Record your readings when you have added sufficient soap solution to produce, after shaking well, a lather covering the surface of the liquid for one minute. It will be found that much less soap is required to produce a lather in the first and second flasks than in the third.

We therefore find that *temporarily hard water* is *softened* (*a*) by adding *lime*, (*b*) by *boiling.* Hence the term '*temporary*' hardness. In both cases *carbonic acid gas is removed from the solution* and thus the *precipitate of calcium carbonate first formed* appears again as a *solid* which has no action on soap. Notice that it is the calcium compound in solution and not the carbonic acid gas which causes the water to be hard.

77. Permanently Hard Water is not softened by either addition of lime or by boiling. Calcium bi-carbonate is decomposed by heat; it is *unstable.* All other calcium salts such as the sulphate and chloride are not affected if their solution is boiled. Their solution may be softened (*a*) by adding washing soda (sodium carbonate); (*b*) by adding soap (sodium stearate); in both cases an insoluble calcium compound is formed.

***Exp.** i. The solution in the generating flask (§ 75) contains *calcium chloride.* To a few drops of this solution add *washing soda* (sodium carbonate)

solution. A white precipitate of solid calcium carbonate is formed[1]. Filter and test (with soap solution) whether the filtrate is 'soft.' Pour a few drops of acid on the *residue* on the filter paper and note that it is a carbonate.

***Exp.** ii. *Boil* one of two small equal quantities of diluted *calcium chloride* solution and notice that the same amount of *soap* solution is required to obtain a permanent lather in each case.

78. Double Decomposition. Reactions of this class are common in chemistry. Both substances decompose and interact. Thus in Exp. i

(Soluble) Calcium Chloride +
„ Sodium Carbonate
} yield { (Soluble) Sodium Chloride +
(Insoluble) *Calcium Carbonate* ↓ precipitated.

In Exp. ii.

(Soluble) Calcium Chloride +
„ Sodium Stearate (soap)
} yield { (Soluble) Sodium Chloride +
(Insoluble) *Calcium Stearate* ↓ precipitated.

79. Calcium Chloride *obtained from the solution of Marble in Hydrochloric Acid.* When the liquid in the generating flask (§ 75) no longer causes an effervescence with marble, filter about 15 c.c. into a porcelain dish and evaporate to dryness. The white solid is calcium chloride. It is chiefly used as a **drying reagent.**

***Exp.** Weigh the dish containing dry calcium chloride and put it aside for a few hours and weigh again. If the air is very damp the calcium chloride may have absorbed so much moisture as to be in a state of solution. A substance behaving in this way is said to **deliquesce.**

80. Important Revision Experiments. The qualitative composition of carbon dioxide and calcium carbonate may be shown in the following series of experiments, both by analysis and synthesis.

[1] In permanently hard natural waters, *calcium sulphate* (gypsum) is the salt usually present.

(i) *Pass carbon dioxide over* filings of the metallic element *calcium* heated in a bulb tube (Fig. 58). The gas must be freed from hydrochloric acid by passing it through water. The metal burns with a beautiful orange-red glow, forming *calcium oxide* (quicklime); simultaneously *carbon* is liberated. The calcium oxide on cooling combines with the excess of carbon dioxide forming *calcium carbonate* (chalk).

(ii) Break the bulb tube and, having ground up the black mass, treat it with hydrochloric acid in a test tube. The carbonic acid gas, liberated from the calcium carbonate, is identified by passing it through lime water (Fig. 59).

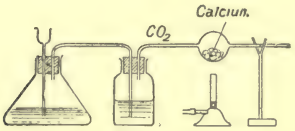


Fig. 58.

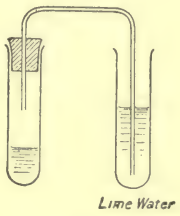


Fig. 59.

(iii) The black substance suspended in the hydrochloric acid solution proves to be *carbon*; filtered, dried and heated in a current of air, it yields carbon dioxide.

The reactions are shown by the following equations:

(1) Calcium + carbon dioxide = calcium oxide + carbon.
Calcium oxide + carbon dioxide = calcium carbonate.

(2) Calcium carbonate + hydrogen chloride = calcium chloride + water + carbon dioxide.

(3) Carbon + oxygen = carbon dioxide.

81. Natural Waters.

A. **Rain Water** is the purest form of natural water since it is formed by evaporation and condensation. On evaporating to dryness there is practically no residue but it contains in

solution the gases present in the atmosphere, viz. oxygen, nitrogen, carbonic acid gas.

B. **Spring Water.** If we evaporate well-water to dryness the solid mineral matter obtained depends on the rock or soil of the country through which the rain water has percolated before collecting in the well. Decaying animal and vegetable matter on the surface impart carbonic acid to the water and thus give it the power to dissolve chalk and limestone. The limestone caves of Yorkshire are formed by water charged with carbonic acid gas and the spring water from a limestone or a chalk district will possess much *temporary hardness* (§ 76) due to the presence of *calcium bi-carbonate.* The *sulphates* of *calcium* and *magnesium* cause *permanent hardness.* Common salt (sodium chloride) is generally present in small quantities.

C. **Mineral Waters.** The waters of Buxton, Bath, Harrogate and other mineral spas contain salts and gases which possess medicinal properties. Common salt, compounds of soda, potash, calcium, magnesium and iron (chalybeate) as well as excess of the commonly occurring gases and sulphuretted hydrogen may be present in mineral waters.

D. **River Water** contains (*a*) solids *in solution* brought down from the land; and (*b*) solids *in suspension*—mud, consisting of mineral matter and decaying animal and vegetable remains.

E. **Sea Water** contains more solids in solution than spring and river waters, because the rivers are constantly bringing down material from the land while only pure water is leaving the ocean by evaporation. The Indian Ocean where much evaporation is going on is more saline than the north Atlantic. The Baltic is only brackish on account of the large supply of fresh water from numerous rivers. Sodium chloride (salt) is the chief solid in solution but in addition there are bromides, sulphates and carbonates of potassium, magnesium

and calcium. In the Indian Ocean the total of solids in solution is about 3·5 per cent.

82. Spring water contains air in solution.

***Exp.** i. Fill a 250 c.c. flask completely with water. A rubber cork, with one hole, is fitted to a long delivery tube so that the end of the tube is flush with the cork. The tube is now filled with water and the cork inserted in the flask: the apparatus is then connected to a pneumatic trough (Fig. 60). Boil the water and collect any gases driven off in a narrow tube.

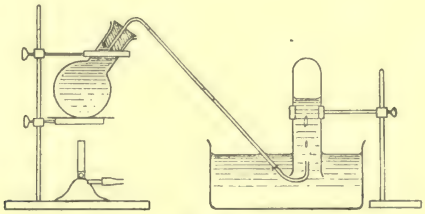


Fig. 60.

Exp. ii. **To analyse the air collected**—(1) push up into the tube a small stick of yellow *phosphorus* attached to a wire. [**Danger.**] *Oxygen* will be slowly absorbed. (2) Replace the phosphorus with a stick of *caustic soda, carbonic acid gas* will be absorbed. (3) The remaining gas will be nitrogen. An actual experiment gave 25 % oxygen, 25 % carbonic acid gas, 50 % nitrogen.

Questions on Chapter XII.

1. Explain clearly what is the connection between chalk, lime and carbon dioxide. (O. J.)

2. Describe experiment by which the percentage weight of carbon dioxide in chalk may be determined (*a*) by heating, (*b*) by use of acid.

3. 0·74 gram of chalk on being heated yielded ·415 gm. of quicklime. Calculate percentage composition of chalk.

4. How may the volume of carbonic acid gas in 100 gm. of calc-spar be determined?

5. In what way does magnesia resemble quicklime?

6. If the density of carbon dioxide is 22 what is the weight of 5 litres of the gas?

7. How would you obtain from chalk a small specimen of the carbon which it contains? How would you prove that pure calcium carbonate contains lime and carbon dioxide and nothing else? (O. J.)

8. When carbon dioxide is passed into lime water a milky precipitate is obtained which disappears as more gas passes in. On boiling the clear solution the precipitate reappears. How do you account for this?

9. How would you prepare carbon dioxide and identify the gas?

10. What is meant by 'temporarily hard' water? How may it be softened?

11. To what is permanent hardness due? Why does hard water cause soap to be wasted?

12. How may the 'hardness' of water be determined?

13. Describe what happens when carbon dioxide is passed over heated calcium.

14. How would you determine the amount of air dissolved in a litre of water? In what way does the air differ from ordinary air?

15. Water is stated to be a powerful 'solvent.' What does this mean? How does the fact help to account for the properties of spring waters? (O. J.)

16. What substances other than water are usually found in river water? From what sources are they derived? Why does spring water often become turbid when boiled for some time? (O. J.)

17. From the results of analysis of the air dissolved in spring water (§ 82) what additional reason have you for stating that air is a mixture?

18. Separate quantities of 20 c.c. of distilled water, tap-water, rain-water and distilled water through which carbon dioxide has been bubbled for some time, have soap solution added till a lather is obtained. Should the amount of soap solution required be the same in each case? Give reasons.

CHAPTER XIII.

CARBON AND ITS TWO OXIDES. EFFECT OF PLANT AND ANIMAL LIFE ON THE ATMOSPHERE.

83. Allotropic forms of Carbon.

***Exp.** i. Heat about two grams of *sugar* in a test tube keeping the tube nearly horizontal so that the water liberated does not run down on the hot glass. Notice that an inflammable gas is given off as well as steam and that finally a black substance, which is almost pure *carbon*, remains.

***Exp.** ii. Transfer the carbon to a piece of hard glass tubing connected with a wash bottle containing lime water, and leading to an aspirator (see Fig. 54). Burn the carbon in a current of air. Notice that chalk (calcium carbonate) is precipitated in the lime water and that finally no residue is left. Evidently the carbon burns in oxygen to form **carbonic acid gas** (carbon dioxide). (See Chap. XII.)

Exp. iii. **To find the composition of carbon dioxide by weight,** about 0·2 gram of *carbon* is accurately *weighed* and heated in a long tube containing pure dry oxide of copper (red-hot) and a slow current of dry

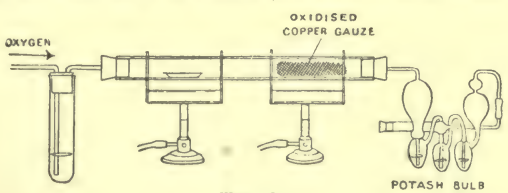


Fig. 61.

oxygen, free from carbon dioxide, is passed through (Fig. 61). The resulting *carbon dioxide* is caught in *weighed* bulbs containing strong *caustic potash* solution to which a calcium chloride tube is attached to retain any moisture which might be lost by the solution. By this method 3 parts by weight of carbon yield, when burnt in oxygen, 11 parts by weight of carbon dioxide, therefore, by *weight*, 3 *parts of carbon combine with* 8 *parts of oxygen.*

Sugar-charcoal, graphite (plumbago, black-lead) and **diamond** all burn completely in oxygen and form carbon dioxide, in the proportions obtained in Exp. iii. We therefore must conclude that these substances are different forms of the same element carbon. Such forms of an element are said to be **allotropic,** *i.e.* the same from a *chemical* standpoint but differing in *physical* properties. Sulphur and phosphorus also exhibit allotropic forms[1].

84. The Law of Fixed Proportions. Exp. iii (§ 83) is an instance of a method of testing whether a substance is a chemical compound or not, and exemplifies one of the fundamental laws of chemistry. In this case pure carbon dioxide, no matter how it is prepared, always contains carbon and oxygen united together in the proportions 3 to 8 by weight.

The same chemical compound always contains the same elements united together in fixed proportions by weight.

This law is also called the *Law of Constant* or *Definite Proportions.*

85. Carbonic Oxide. If a slow current of carbon di-oxide is passed through a hard glass tube containing charcoal

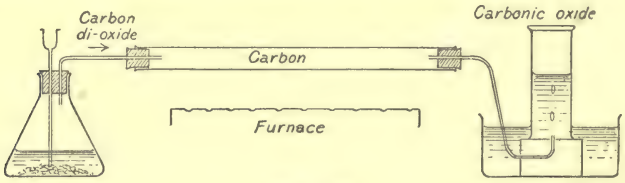


Fig. 62.

heated red-hot (Fig. 62), an inflammable gas issues from the tube. It can be proved that it contains carbon and oxygen united in the proportions of 3 to 4 by weight. The gas is called

[1] *Charcoal, coke, anthracite coal, lamp-black* and *soot* are other forms of nearly pure carbon but they can hardly be classed as allotropic forms.

carbonic oxide (carbon monoxide). Its density is 14 (hydrogen = 1) and it burns with a blue flame in oxygen forming carbon dioxide. It is *poisonous*. Its *composition by weight* may be obtained by passing a *known weight* of carbonic oxide (free from carbon dioxide and water vapour) through a long glass tube containing red-hot oxide of copper. The latter substance is *reduced* to copper and the *carbonic oxide is oxidized to carbon dioxide* which may be caught in *weighed* potash bulbs. (See Fig. 61, Exp. iii, § 83.)

86. The Law of Multiple Proportions. Compare the composition by *weight* of these two oxides of carbon:

	Carbon	Oxygen
Carbon *di*oxide (carbonic acid gas)	3	8 parts by weight
Carbon *mon*oxide (carbonic oxide)	3	4 „ „

If two elements (such as carbon and oxygen) *combine to form two or more compounds* (such as carbon dioxide and carbonic oxide) *and the weight of one element* (say carbon) *is kept constant* (here = 3) *then the weights of the other element* (oxygen) *are in simple multiple proportion* (here 8 : 4 = 2 : 1).

87. The Chemistry of a Red-hot Fire. We have noticed that on the top of a glowing, red-hot coal, coke or charcoal fire, there is often playing a *pale blue flame*. It is *carbonic oxide* burning with the oxygen of the air drawn in from the room. The diagram (Fig. 63) gives us an idea of the changes.

At *AA* carbon combines with oxygen to form carbon dioxide.

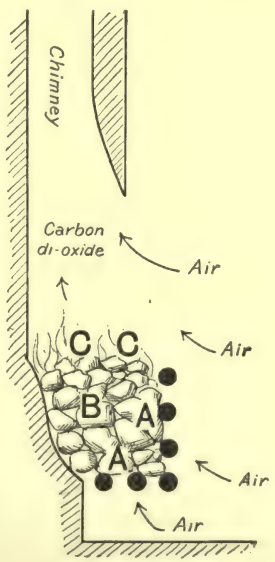


Fig. 63.

At *B* carbon dioxide loses half its oxygen to form carbon monoxide (carbonic oxide) and at the same time this oxygen that it has lost combines with carbon to form more carbonic oxide.

At *CC* carbonic oxide burns with oxygen to form carbon dioxide. (Cf. reduction of carbon dioxide to carbonic oxide, § 85).

88. Burning, Breathing, Fermentation and Decay.

If we confine our attention to the various forms of matter which contain **carbon** (usually called *organic* matter), we find that in the above processes the carbon is oxidized to **carbon dioxide.**

***Exp.** i. *To show that* **carbon dioxide** *and water* are formed when a **candle is burned,** lower a lighted candle fixed to a deflagrating spoon into a jar of air and close the top of jar with a piece of cardboard. Moisture condenses on the sides and on shaking the gas remaining, after the candle flame dies away, with *lime water*, a precipitate of calcium carbonate is formed.

The experiment may be varied by lowering a jet of burning *coal gas*, or a deflagrating spoon containing ignited *methylated spirit*, *turpentine*, *naphtha petroleum*, *benzene*, &c., into the jar of air.

***Exp.** ii. **Breathe** into lime water, through a glass tube; *or* set up two wash bottles (Fig. 64) and breathe in and out through the lime water and note the results.

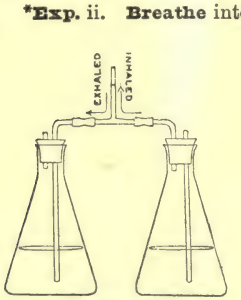


Fig. 64.

By using the pneumatic trough, collect a jar of breath, deeply expired, and test whether a taper is extinguished in the air breathed out.

***Exp.** iii. Mix together a solution of *sugar* and a little *yeast* in a flask connected by a tube to a test tube containing lime water. Allow the mixture to **ferment** in a warm place for some hours and note the liberation of carbon dioxide.

***Exp.** iv. **Oxidation of organic substances.** To test whether a substance is **organic,** *i.e.* a **carbon** compound, the substance is mixed with dry *oxide of copper* and placed in a hard glass test tube: a layer of copper oxide is then added and the contents of the tube are heated and the gas liberated is passed through lime water. The layer of copper oxide is strongly heated *first* (Fig. 65) to ensure the oxidation of any

vapours which may be driven off. If *water* is also formed, the presence of *hydrogen* in the organic substance is proved. In this way, test whether

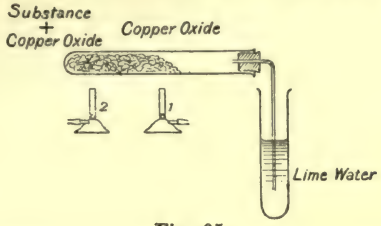


Fig. 65.

naphthalene, *starch*, *toasted bread*, and *soap powder* each contains both carbon and hydrogen.

Nearly all **fuels** and **food-stuffs** contain both carbon and hydrogen, consequently in ordinary cases of burning and in the slow combustion of our bodies both carbon dioxide and water are formed. The blood carries oxygen to all parts of our tissues where chemical changes go on and at the same time heat is produced. In a healthy human being the normal temperature is 98·4° F. When 'combustion' proceeds too rapidly, or when heat does not leave our bodies by the usual processes of perspiration and evaporation, our temperature rises and we become *feverish*.

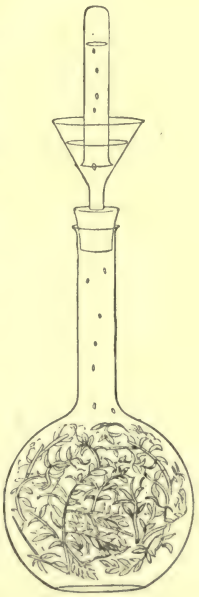


Fig. 66.

89. Mutual interchange of carbon and oxygen between animals and plants. Even the flesh-eating animals such as lions, tigers, eagles, vultures and hawks, obtain their food ultimately from plant life. The bullock eats grass, the tiger kills the bullock and eats him. The grass contains carbon, the bullock's flesh contains carbon, and both

tiger and bullock breathe out carbonic acid gas into the air. The grass must obtain its carbon from the air, for we can grow grass in soil which contains no carbon. Hence we see that there is a constant interchange of carbon between plants and animals: animals either directly or indirectly live on plants; the carbon of the food which animals eat oxidizes in their bodies

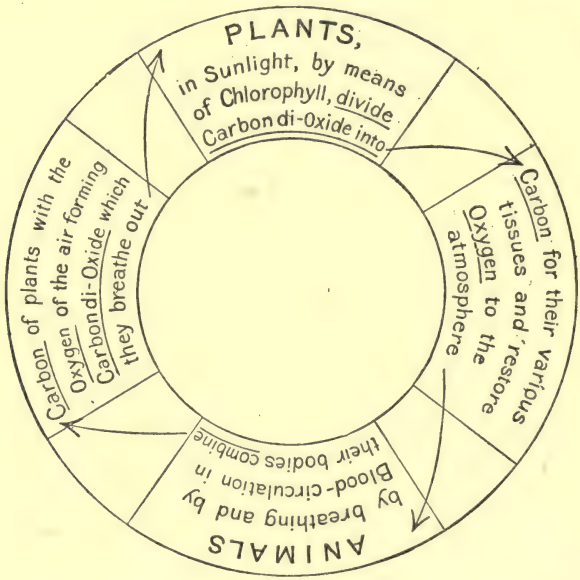

PLANTS,
in Sunlight, by means of Chlorophyll, divide Carbon di-Oxide into
Carbon for their various tissues and restore Oxygen to the atmosphere
ANIMALS
by breathing and by Blood-circulation in their bodies combine
Carbon of plants with the Oxygen of the air forming Carbon di-Oxide which they breathe out

Fig. 67.

and passes out into the atmosphere as carbon dioxide: plants on the other hand take carbon from the carbon dioxide and restore the oxygen to the atmosphere. If this cycle of changes did not take place, it is evident that the air would soon be deprived of its oxygen and the result of the processes of burning, breathing and decay would be that carbon dioxide would accumulate until the air was unfit for animal life.

Chlorophyll, the green colouring matter in plants, has the property of dividing carbon dioxide under the action of sunlight; the **carbon** is kept to build up the structure of the plant and the oxygen is restored to the air. The process is a complicated one, but its result may be easily shown by exposing water plants such as duckweed or watercress to sunlight in a jar containing water charged with carbonic acid gas. Bubbles appear on the leaves and these bubbles accumulate until sufficient gas is obtained to show the test for oxygen. In Fig. 66 green water weeds are put into a large flask filled with 'soda' water and exposed to sunlight. The gas liberated (oxygen) is caught in a test tube and afterwards tested with a glowing splinter.

This cycle of changes is shown graphically in Fig. 67.

Pure country air contains 4 parts of carbon dioxide in 10,000

Questions on Chapter XIII.

1. What do you mean by 'allotropic forms' of carbon? Mention any you know and state for what purposes they are used.

2. Describe an experiment to prove that carbon dioxide contains three parts by weight of carbon combined with eight parts of oxygen.

3. How could you distinguish between carbon dioxide and monoxide?

4. How may the composition by weight of carbon monoxide be found?

5. State the Law of Multiple Proportions and give some proof of its truth.

6. What substances are produced when a candle burns in air? How would you identify them?

7. By what means could you show that soap is an organic substance?

8. State what is meant by 'combustion' and 'respiration.' In what way do the two processes resemble one another?

9. What portions of its food does a plant obtain from the air and under what circumstances can it absorb them? (O. J.)

10. Given a mixture of carbon dioxide and oxygen, what means would you employ to obtain some pure oxygen from it?

11. How could you distinguish carbon dioxide from nitrogen? How could you prove that carbon dioxide contains its own volume of oxygen? (O. J.)

CHAPTER XIV.

SALT, HYDROCHLORIC ACID AND CHLORINE.

90. Rock Salt occurs in many parts of the world where ancient inland seas have dried up. From some cause or another, a layer of impervious clay has been deposited above the salt, thus protecting it from being dissolved by water which might otherwise have percolated through it. In Cheshire, beds several yards in thickness of almost pure salt are found at depths varying from 120 to 400 yards. In some places mines are sunk, but as a rule the salt is obtained by boring two holes down one of which water sinks into the salt-bearing strata, while up the other a saturated solution (*brine*) is pumped. This *brine* is evaporated until crystals form.

Sodium Chloride (common salt) is a compound consisting of two elements, *sodium* (metal) and *chlorine* (non-metal) united in the proportion by weight 23 : 35·5 (see § 22). Its curve of solubility is given in § 28. Its crystals are transparent cubes (m.p. 815° C.).

Preparation. Sodium chloride is prepared chemically by :

(1) Direct union of sodium metal and gaseous chlorine (§ 23).
(2) By acting on metallic sodium with hydrochloric acid (§ 93).
(3) „ sodium oxide „ „ (§ 93).
(4) „ sodium hydroxide „ „ (§ 59).
(5) „ sodium carbonate „ „ (§ 73).

This summary shows the **General Methods for the preparation of salts.** If the salt required is 'neutral,' the

solution must be carefully tested with a suitable indicator and afterwards crystals obtained by evaporation (see §§ 58 and 59).

91. Hydrochloric Acid (Spirits of Salt, or Hydrogen Chloride) is a colourless gas, fuming in moist air and possessing an irritating odour. It is exceedingly soluble in water; one volume of water dissolves more than 500 volumes of the gas at 0° C.

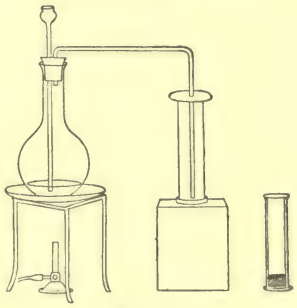


Fig. 68.

Preparation. ***Exp.** Take about 30 c.c. of water in a beaker and into it *slowly* pour 40 c.c. of strong *sulphuric acid*; stir well with a glass rod. Place about 30 grams of *rock salt* or fused common *salt* in a 250 c.c. flask fitted with a thistle funnel and tube. Set up the apparatus shown in Fig. 68, and pour the diluted sulphuric acid down the funnel on to the salt. Heat gently. Collect a cylinder of the gas by *downward* displacement (Density = 18) and show that a taper will not burn in it.

The reaction is represented by the following equation:

$$\left.\begin{matrix}\text{Sodium Chloride} + \\ \text{Hydrogen Sulphate}\end{matrix}\right\} = \left\{\begin{matrix}\text{Sodium Sulphate} + \\ \text{Hydrogen chloride}\end{matrix}\right.$$

92. To show the solubility of Hydrochloric Acid Gas.

***Exp.** i. Fill a dry round-bottomed flask with the gas by 'downward' displacement. Fit a rubber cork, through which a tube drawn out to a jet passes into the flask and invert the tube in a beaker of water coloured *blue* with litmus. Cool the flask by pouring a little ether over it. The water runs up the tube and dissolves the hydrochloric acid gas, thus producing a partial vacuum. Record all you observe and give your reasons for what takes place.

The solubility may be similarly shown by filling a tube with the gas, covering the end with the thumb and inverting it in water. On removing the thumb water fills the tube.

***Exp.** ii. Make a solution of the gas by passing it into water through a wide funnel just dipping below the surface of the water (Fig. 70). Weigh

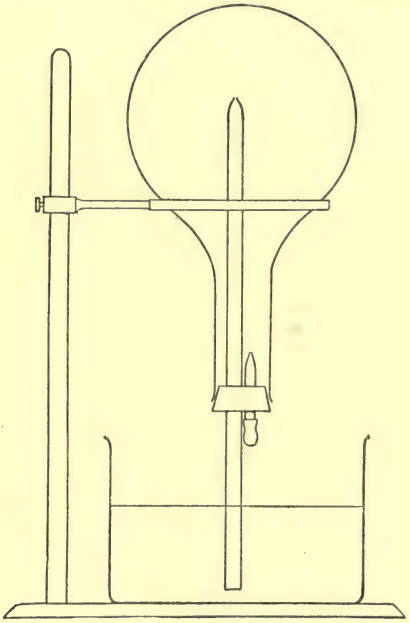


Fig. 69.

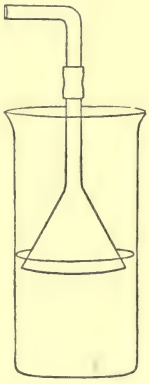


Fig. 70.

the water before and after dissolving the gas. The density of concentrated hydrochloric acid = 1·2.

For experiments with the acid refer back to § 55.

Tests. Hydrochloric acid is recognized by the following:

(1) has effect on litmus and other indicators,

(2) fumes in air and has a characteristic odour,

(3) is very soluble in water,

(4) forms dense clouds of *ammonium chloride* (solid) with ammonia gas (§ 57),

(5) forms with *silver nitrate* solution a white curdy precipitate, which turns violet in sunlight,

(6) on passing an electric current through its solution for some hours by carbon electrodes (§ 120) equal volumes of *hydrogen* and of *chlorine* are finally evolved.

93. Composition of Hydrochloric Acid Gas.

1. By **synthesis.** If one volume of *hydrogen* and one volume of *chlorine* (prepared by electrolysis) are exposed in a glass bulb to sunlight, they combine with explosion; but if

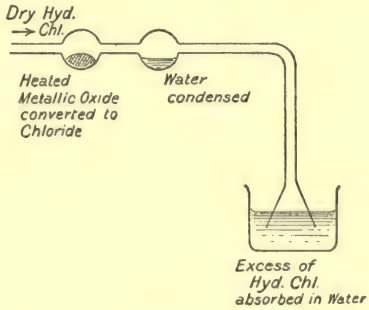


Fig. 71.

exposed to much less brilliant daylight they combine gradually and form two volumes of hydrochloric acid gas.

2. If a jet of *hydrogen* is burned in *chlorine* gas acid fumes of hydrogen chloride are produced (§ 95).

3. By **analysis.** If dry hydrochloric acid gas is passed over heated *sodium*, *calcium*, *zinc* or *magnesium*, *hydrogen* gas is liberated and the chloride of the metal is formed.

4. **Exp.** If dry hydrogen chloride gas is passed over the warmed *oxide* of either *lead* (litharge), *zinc*, *calcium*, or *magnesium*, *water* (oxide of hydrogen) is formed and may be condensed in a U or bulb tube (Fig. 71), and the *chloride* of the metal is

formed at the same time. This experiment demonstrates that **when an acid reacts with a base to form a salt water is also produced,** *e.g.*,

$$\left.\begin{matrix}\text{Zinc Oxide} + \\ \text{Hydrogen Chloride}\end{matrix}\right\} = \left\{\begin{matrix}\text{Zinc Chloride} + \\ \text{Hydrogen Oxide (water)}\end{matrix}\right.$$

5. **Exp.** If dry hydrogen chloride gas is passed over *red lead*, a *per*oxide[1] of lead (or over pyrolusite, the *per*oxide[1] of manganese), not only are the metallic chloride (a salt) and water

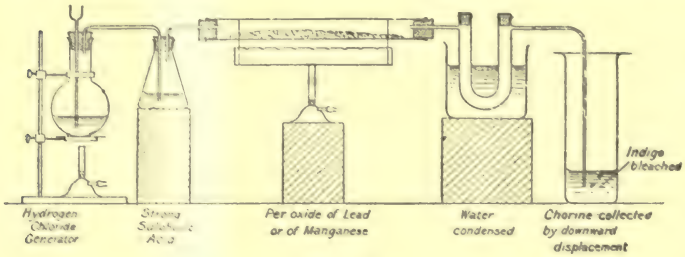


Fig. 72.

formed, but in addition **Chlorine** gas is set free, because there is a surplus of *oxygen* available to combine with the hydrogen of the acid which is in excess and thus form more water and release chlorine.

$$\left.\begin{matrix}\text{Hydrogen Chloride} + \\ \text{'Oxide of Lead Oxide[1]'} + \\ \text{Hydrogen Chloride}\end{matrix}\right\} = \left\{\begin{matrix}\text{Water} + \text{Chlorine} \uparrow + \\ \text{Water} + \text{Lead Chloride}\end{matrix}\right.$$

94. Preparation of Chlorine. The last two experiments exhibit the difference in action of (*a*) a metallic oxide, (*b*) a

[1] A *peroxide* may be considered to be an oxide still further oxidized or an 'oxide of an oxide.'

metallic *per*oxide on hydrogen chloride (hydrochloric acid). The peroxide has been written 'oxide of the metallic oxide' to indicate that there is more oxygen available than is necessary to form a salt and water according to the usual action of an acid on a base. *The additional oxygen liberates chlorine from hydrochloric acid*, although free oxygen gas (in air for instance) has no effect on the acid under ordinary conditions. If however a mixture of air and hydrochloric acid gas is passed over heated bricks previously soaked in a solution of *chloride of copper*, which performs the part of a catalytic agent (§ 47), the oxygen of the air combines with the hydrogen of the acid and chlorine is set free (Deacon's Process).

Hydrogen Chloride + Oxygen = Water + Chlorine.

Exp. The usual **method of preparing Chlorine** in the laboratory is to cover lumps of *pyrolusite* (peroxide of manganese) with strong *hydrochloric acid* in the apparatus shown in Fig. 73. Apply gentle heat. The gas is passed through a wash bottle containing water to remove acid which may be carried over. Collect the gas by *downward* displacement.

Another excellent method is to drop strong hydrochloric acid slowly from a tap-funnel on to the strong oxidizing agent potassium permanganate (crystals). (See Fig. 17.)

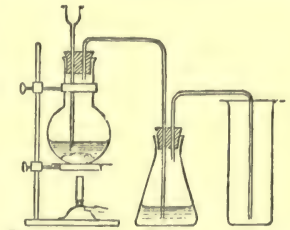
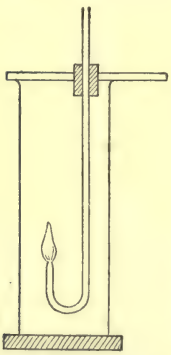


Fig. 73.

Properties of Chlorine: Soluble in water; colour, yellowish green; density = 35·5 (hydrogen = 1); smell, very irritating; injurious (should be prepared in *draught-place*); combines directly with many elements; *bleaches*.

95. Experiments with Chlorine. Having collected jars of chlorine and having covered them with greased plates, perform the following experiments:

Experiment.	**Observations.**
1. (*a*) Put a piece of dried turkey-red cloth and a dried envelope which has passed through the post in a jar and allow to stand for a few minutes.	(*a*) Very little change—dry chlorine does not bleach.
(*b*) Next add a little water and shake.	(*b*) The cloth, writing ink and stamp are **bleached,** but the postmark is not. Printers' ink contains free *carbon* which does not readily react with chlorine.
2. Drop powdered antimony into the second jar.	The metal burns and chloride of antimony is formed.
3. Pour a few drops of *turpentine* (containing carbon and hydrogen) on a filter paper, previously warmed, and drop it into the third jar.	Charring occurs, possibly with flame: fumes of hydrochloric acid are formed and carbon is set free.
4. Burn coal-gas or a taper (containing carbon and hydrogen) in chlorine.	Carbon liberated, hydrogen chloride (acid) formed.
5. Burn a jet of hydrogen in chlorine. (Fig. 74.)	Hydrochloric acid fumes are formed.



Jet of Hydrogen or Coal Gas burning in Chlorine

Fig. 74.

Bleaching Powder. The chief use of chlorine is for *bleaching.* If chlorine is passed over slaked lime, the two substances combine to form *bleaching powder,* which yields up its chlorine again on being treated with dilute acid.

***Exp.** Mix a few grams of bleaching powder with water and soak a piece of turkey-red cloth in the thin paste so obtained. Transfer the cloth to a beaker containing dilute hydrochloric acid. Repeat the operation until the cloth is bleached. Finally wash the cloth well with water to remove all acid.

Questions on Chapter XIV.

1. By what methods may salt be prepared chemically?

2. Sketch the apparatus required for preparing hydrochloric acid gas and describe the experiment in detail. What reason is there for or against collecting the gas over water?

3. By what tests would you identify hydrochloric acid gas?

4. Describe an experiment by which the composition by volume of hydrochloric acid gas may be determined.

5. How may chlorine be prepared? State its chief properties.

6. Light sometimes causes chemical action. Give an instance of this.

7. Describe and explain what happens when dry hydrochloric acid gas is passed over lead peroxide.

8. How would you prove that hydrochloric acid gas is extremely soluble in water?

9. Mention any bleaching agents you know and state on what their action depends.

10. What will take place when a dilute solution of hydrochloric acid is

(*a*) added to a well-water in a limestone district,

(*b*) poured upon zinc metal at the ordinary temperature,

(*c*) warmed with powdered manganese dioxide? (O. J.)

CHAPTER XV.

SULPHUR AND ITS OXIDES, SULPHURIC ACID, AND HYDROGEN SULPHIDE.

96. Sulphur or brimstone occurs in the region of volcanoes, some of which are at present active. Sicily, Japan, Chili, Peru and Nevada possess large sulphur deposits and mines. The crude or native sulphur is heated in retorts and the vapour is condensed. Sulphur melts at 114° C. and boils at 444° C. Its density is about twice that of water. It is soluble in carbon bisulphide (§ 25).

Effect of heat on sulphur. We have already found that sulphur readily melts (§ 18), forming a yellow liquid which darkens and becomes like very thick treacle at about 180° C. On heating further it becomes more liquid again until it reaches the boiling point, when it gives off dark red vapours. We noticed that these vapours condense partly to liquid sulphur and partly to 'snow' or 'flowers of sulphur,' a yellow powder. If sulphur is gradually raised to a temperature of about 130° and the clear yellow liquid is allowed to cool, transparent crystals (needles) are obtained at 120° C. These crystals become opaque in a few hours. If we pour boiling sulphur into cold water, *plastic sulphur*, having the consistency of soft india-rubber, is formed. This plastic variety is *unstable* and turns to the opaque variety in a few days. These three forms are all chemically pure sulphur, but they differ in physical properties, *e.g.* density, transparency, hardness and solubility in carbon bisulphide. They are **allotropic forms** (see § 83).

Metallic Sulphides. We have seen that iron and copper burn in sulphur vapour (§ 20) and form black sulphides of the metal. Beautiful cubical crystals of galena (sulphide of lead), iron pyrites and copper pyrites (sulphides of iron and copper) are found in nature and should be examined and described. Blende (sulphide of zinc) and cinnabar (sulphide of mercury) are ores which readily yield their metals.

97. Sulphur dioxide. We are well acquainted with the smell of burning sulphur and have already found that when sulphur combines with oxygen a colourless gas, acid and suffocating, is formed. This gas is *sulphur dioxide* or *sulphurous anhydride*. It is also produced when *iron pyrites* is roasted in a current of air. We discovered too, that it was liberated when copper, zinc and some other metals were heated with *strong* sulphuric acid (§ 53).

Preparation of Sulphur dioxide.

Exp. The experiment should be performed in the draught-place. In a flask of 250 c.c. capacity fitted with cork, thistle funnel and 'downward delivery' tube, place about 15 grams of copper turnings. Fit up the apparatus (Fig. 75). Pour about 40 c.c. of strong sulphuric acid down the funnel and warm gently. Collect three jars of the gas by 'downward displacement' and cover the jars with glass plates. Then pass the gas into water by means of a funnel just dipping beneath the surface of the liquid (Fig. 70).

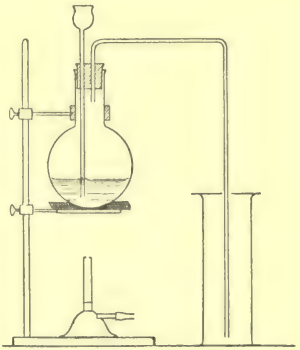


Fig. 75.

Experiments with Sulphur dioxide.

Exp. i. Pour blue litmus solution into a jar of the gas; cover the jar with the plate and shake well. The gas dissolves, forming an acid.

Exp. ii. Lower a lighted taper into a jar of sulphur dioxide. Relight the taper and pour the gas out of the jar over the flame. Note that in both cases the taper is extinguished.

Exp. iii. Show the slow bleaching action of the gas by putting coloured flowers into the third jar.

Exp. iv. Using the solution you have prepared—

(*a*) find whether it remains acid after being boiled for some minutes.

(*b*) try the effect of adding the solution to sodium permanganate solution (Condy's fluid). Is the smell removed? Is the solution still acid?

(*c*) repeat (*b*), substituting potassium bichromate solution.

Properties of Sulphur dioxide. Its density = 32 (hydrogen = 1); at 15° C. one vol. of water dissolves about 50 vols. of gas. Cooled to − 10° C. it condenses to a watery liquid which may be kept as a liquid under pressure in siphons. It combines with water to form *sulphurous acid* and with bases to form *metallic sulphites* (salts).

Composition. If a weighed quantity of sulphur is burned in a current of oxygen and the sulphur dioxide is caught and weighed in weighed tubes containing alkali, the sulphur is found to have combined with an equal weight of oxygen.

Preparation of Copper Sulphate and recovery of Copper.

Exp. Allow the liquid in the generating flask to cool; pour away the discoloured acid from the grey sediment (anhydrous copper sulphate) and dissolve the latter in a little water; filter and evaporate to obtain crystals. Boil some of the mother liquor with iron filings and note the displacement of copper.

98. Sulphur trioxide or Sulphuric Anhydride is a white solid, crystallizing out from the gaseous state in silky needles which melt at 15° C. and which readily combine with water to form *sulphuric acid* (see § 99, Exp. iii).

99. Sulphuric Acid (Oil of Vitriol) **or Hydrogen Sulphate** has already been described in § 52. It was first obtained by heating green vitriol (ferrous sulphate) crystals very strongly.

***Exp.** i. Heat finely ground crystals of green vitriol until the water of crystallization has been expelled. Then transfer the 'anhydrous' (lit.

without water) residue to a test tube and heat until a red powder, a red oxide of iron (jeweller's rouge), is obtained. At the same time notice that strongly acid fumes, which may be condensed to sulphuric acid, are given off.

Exp. ii. The following experiment illustrates the commercial method of making sulphuric acid. Sulphur dioxide, oxygen (of the air) and steam are passed into a large **leaden chamber** which gives its name to this process. Nitric acid fumes are added which act as a catalytic agent, causing combination of the first three named substances. Into a large bolt-head (Fig. 76)

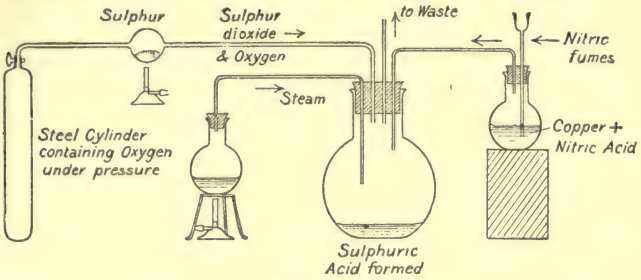


Fig. 76.

pass (1) oxygen and sulphur dioxide produced by burning sulphur in excess of oxygen, (2) steam, (3) the nitric fumes formed by pouring nitric acid on copper. Sulphuric acid condenses in the bolt-head, which corresponds to the 'leaden chamber,' and any uncombined gases are carried away to the draught-place.

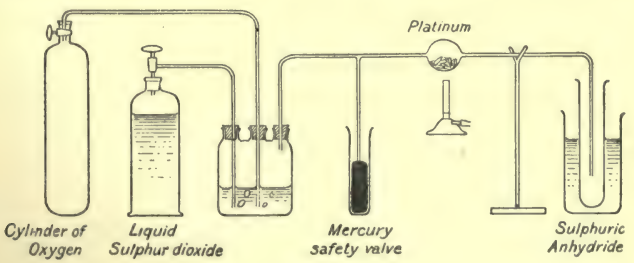


Fig. 77.

Exp. iii. **The contact method.** A more modern process of preparing sulphuric acid is as follows: Dry oxygen and sulphur dioxide, although usually inert towards each other, combine readily when they are in *contact* with heated finely divided *platinum*. Through a wash bottle (Fig. 77)

containing strong sulphuric acid to dry the gases, sulphur dioxide (from a siphon) and oxygen (from a steel cylinder) are passed in the proportions of 2 to 1 by volume. The mixed gases are conducted over warmed spongy platinum and combine to form the anhydride (lit. 'without water') of sulphuric acid. Silky crystals of the anhydride may be obtained by condensing the fumes in an ice-cooled tube or sulphuric acid may be formed directly by passing the fumes into water through a wide funnel.

100. Test for Sulphuric Acid and a Sulphate.

Apart from the reactions given in § 53, sulphuric acid may be detected by the fact that the only common compound of the metal barium, insoluble in hydrochloric acid, is barium sulphate.

***Exp.** Boil the liquid, in which the presence of sulphuric acid is suspected, with hydrochloric acid, filter if necessary, and add a solution of a barium salt. A white precipitate indicates the presence of barium *sulphate* and therefore of a sulphate in the original solution.

Practical Exercise. Show that sulphurous acid solution is converted to sulphuric acid by treating it with (*a*) strong nitric acid, (*b*) potassium bichromate, (*c*) potassium permanganate.

101. Hydrogen Sulphide (Sulphuretted Hydrogen)

is a gas possessing the smell of rotten eggs. It is slightly soluble in water and occurs in Harrogate and other spa waters. It is *poisonous* and experiments with it should be made in the draught-place. Small quantities are formed when hydrogen gas is passed through the vapour from boiling sulphur, thus showing that it is a compound of the two elements. The usual method of preparation is to act on a metallic sulphide with a mineral acid. By double decomposition (§ 78) the hydrogen of the acid combines with the sulphur of the sulphide while the metal takes the place of the hydrogen in the acid; *e.g.*,

Iron Sulphide + Hydrogen Chloride
= Iron Chloride + Hydrogen Sulphide.

***Exp.** Repeating the experiment of § 20, Exp. iii, heat strongly a mixture of iron turnings with sulphur in a test tube. The iron burns and combines with the sulphur forming sulphide of iron. Pour hydrochloric or sulphuric acid on the sulphide—the disagreeable odour of sulphuretted hydrogen is obvious. The presence of the gas may be detected by the fact

that it blackens filter paper moistened with a solution of a *lead* salt—the sulphide of lead being black.

102. Preparation and Properties of Hydrogen Sulphide. In finding out what metals may be present in ores and other substances, the chemist makes use of this gas to precipitate the sulphides of metals, some of which are soluble in water; others, insoluble in water, are soluble in acids or in alkalies. Hence a *continuous* stream of the gas is often required for the complete precipitation of the sulphide, a process which may occupy some hours.

To prepare a continuous current of the gas. The following methods apply to the preparation of any gas formed by the action of a liquid on a solid; *e.g.*,

Liquid acting on	**Solid** liberates	**Gas**
Hydrochloric acid	Marble	Carbonic acid gas
Dilute sulphuric acid	Zinc	Hydrogen gas
Dilute hydrochloric acid	Iron (ferrous) sulphide	Hydrogen sulphide
Water	Calcium carbide	Acetylene

Fig. 78 shows a type of the apparatus used. A stout test tube has a small hole blown through its closed end. This tube is filled with lumps of the solid (say ferrous sulphide) and is fitted with a cork and tap delivery tube. If the tap is turned off and the tube put into a beaker of dilute acid, the acid will not rise in the tube until air is let out of the test tube by turning on the tap: when the acid (liquid) rises to its own level in the tube, it attacks the sulphide (solid), liberating sulphuretted hydrogen (gas). If the tap is now turned off, gas accumulates in

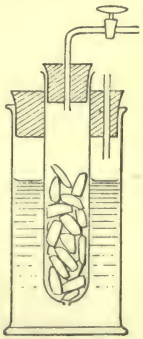


Fig. 78.

the tube and forces the acid out of the hole at the bottom, thus stopping the production of gas.

Kipp's Apparatus. The upper and lower bulbs are connected by a vertical tube and contain the *liquid.* The

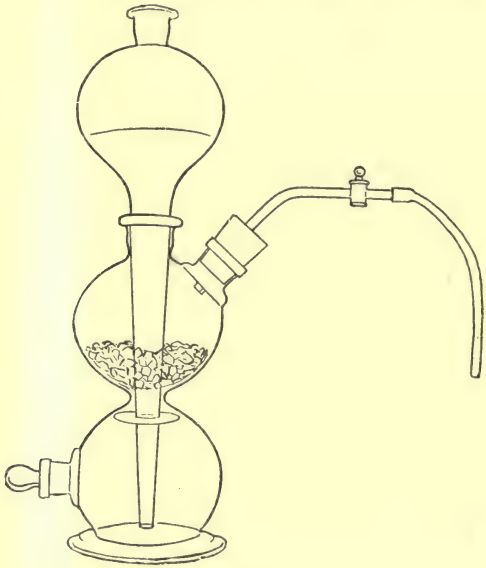


Fig. 79.

middle bulb contains the *solid.* The gas accumulates first in the middle bulb and is delivered by the side tube. In Fig. 79 the apparatus is shown charged.

Properties of Hydrogen Sulphide.

Exp. Prepare hydrogen sulphide, using the above pieces of apparatus or more simply by pouring hydrochloric acid on ferrous sulphide in a flask fitted with cork, funnel and delivery tube (Fig. 56). Collect a few jars of the gas by downward displacement and prepare a solution by bubbling the gas through water.

Show that the gas burns, forming sulphur dioxide and water. If the supply of air is limited by surrounding the flame with a gas cylinder, sulphur is set free.

Show that the gas is slightly acid and that it combines with ammonia, forming a salt (ammonium sulphide) in solution.

Pass the gas into separate solutions of (*a*) copper sulphate, (*b*) ferrous sulphate, to which a little hydrochloric acid has been added. A precipitate of copper sulphide is obtained in (*a*) but no reaction is visible in (*b*). Now make (*b*) alkaline by adding ammonia. A black precipitate of iron sulphide is seen.

Practical Exercise. (1) Devise a method of separating the copper from the iron (as sulphides) in a mixture of the sulphates of the two metals.

(2) Show that on passing the gas through (*a*) nitric acid, (*b*) chlorine water, sulphur is liberated. (See § 104.)

Questions on Chapter XV.

1. By what physical tests could you identify sulphur? For what purposes is sulphur used?

2. Describe all the changes that occur when sulphur is heated.

3. What are the allotropic forms of sulphur? State how each may be obtained and identified.

4. Describe the appearance of galena, iron pyrites, copper pyrites. How could you prove that each contains sulphur?

5. How may sulphur dioxide be prepared? Draw a diagram to show how the gas may be collected. To what uses is this gas put and what takes place when a jar of hydrogen sulphide is opened over a jar of sulphur dioxide? (O. J.)

6. Describe, with sketch of apparatus, how you would determine the composition by weight of sulphur dioxide.

7. State all that occurs when green vitriol is heated.

8. How may sulphur trioxide be prepared? By what means could you identify it?

9. Explain the process employed for the commercial preparation of sulphuric acid.

10. If provided only with iron filings, nitric acid and sulphur, how could you prepare (*a*) sulphuric acid, (*b*) sulphuretted hydrogen? Sketch the apparatus you would employ. (O. J.)

11. Give diagrams of apparatus used for supplying a continuous stream of sulphuretted hydrogen. Explain how they work and try to invent a similar kind of apparatus.

12. Describe sulphuric acid. For what purpose is it used in large quantities? (O. J.)

13. What substances are formed when sulphuretted hydrogen is burnt in air?

14. Four gas jars are supposed to contain sulphuretted hydrogen. Describe four experiments which could be made with them to prove that they do actually contain it. (O. J.)

15. In what way do carbon dioxide and sulphur dioxide resemble one another?

16. What is the effect of passing sulphuretted hydrogen through (*a*) chlorine water, (*b*) a solution of copper sulphate, (*c*) nitric acid, (*d*) a solution of ferrous sulphate containing hydrochloric acid?

17. Suggest some means for proving that sulphuretted hydrogen contains (*a*) hydrogen and (*b*) sulphur but nothing else.

18. By what tests would you identify (*a*) sulphuric acid, (*b*) sulphur dioxide?

CHAPTER XVI.

OXIDATION AND REDUCTION.

103. If we refer to Ch. XI we shall remember that steam (oxide of hydrogen) on being passed over red-hot iron converted the iron to iron oxide, hydrogen gas being set free.

$$\underset{\text{(steam)}}{\text{Oxide of hydrogen}} + \text{iron} = \text{oxide of iron} + \text{hydrogen.}$$

Here steam acts as an **oxidizing agent,** giving its oxygen to iron. On the other hand iron acts as a **reducing agent,** it 'leads back' water to hydrogen again.

We found too that hydrogen, on being passed over hot oxide of iron, combines with the oxygen to form water and turns the oxide back to iron once more.

$$\text{Hydrogen} + \text{oxide of iron} = \underset{\text{(water)}}{\text{oxide of hydrogen}} + \text{iron.}$$

Here oxide of iron is the **oxidizing agent** and is itself **reduced,** *i.e.* 'converted back,' to iron by the hydrogen which is the **reducing agent.**

An **oxidizing agent**[1] is a substance which gives oxygen to another substance, or which reacts in such a way that active oxygen is liberated.

A **reducing agent**[1] is a substance which takes up oxygen from another substance or which reacts in such a way that the other substance loses oxygen.

[1] We shall learn later that these terms have a wider meaning; at present it is sufficient to take them in their simplest sense.

It is evident that in many cases (*e.g.* the examples above) oxidation and reduction go hand in hand. When air turns iron to rust it is simply a case of oxidation due to combination of iron and oxygen. But when hydrogen is passed over red oxide of lead, the hydrogen is oxidized to water and the red lead is reduced to metal. Again when carbon dioxide is passed over red-hot carbon the carbon is oxidized to carbon monoxide, and at the same time the carbon dioxide loses half its oxygen and is reduced to carbon monoxide. (§ 85.)

104. Oxidizing agents. In choosing oxidizing agents we select those rich in oxygen, which readily yield oxygen directly or those which help to liberate it.

Oxygen (either pure or in the air), nitric acid, peroxides of the metals, potassium nitrate, potassium chlorate, potassium bichromate, sodium permanganate (Condy's fluid) and chlorine are common oxidizers.

***Exp.** i. Drop a piece of red-hot charcoal into a few drops of nitric acid in a test tube, keeping your hands away from the tube. The carbon is oxidized to carbon dioxide.

***Exp.** ii. Melt a few crystals of potassium nitrate (nitre) and drop a small piece of charcoal into the tube. Again carbon is oxidized to carbon dioxide.

***Exp.** iii. Try which of the above-named oxidizing agents will convert sulphurous acid to sulphuric acid or sodium sulphite to sodium sulphate (§ 100).

Chlorine acts as an oxidizing agent *in the presence of water* by combining with the hydrogen. The **oxygen** does not come off freely as gas but is in the state of being produced (**'nascent state'**) and is ready to react with, and oxidize, any reducing agent present. Chlorine bleaches because this nascent oxygen destroys the colouring matter.

Oxidizing agents are often good **disinfectants** and **deodorizers.** They convert the carbon and hydrogen of decaying animal and vegetable matter into carbon dioxide and water; they also destroy the microbes.

105. Reducing agents readily combine with oxygen and must be selected according to the conditions of the reaction.

Hydrogen and *carbon monoxide* passed over some of the metallic oxides convert them to metal, at the same time forming water and carbon dioxide respectively.

Coal gas contains hydrogen and compounds of carbon with hydrogen, both of which elements readily combine with oxygen. Hence coal gas may be used to reduce certain metallic oxides.

Carbon alone will also effect reduction.

Sulphur dioxide and *Sulphurous acid* act as reducing agents by combining with oxygen to form sulphuric acid. (§ 100.)

Hydrogen sulphide reduces because its hydrogen combines with oxygen to form water, at the same time sulphur may be set free or it may be oxidized.

***Exp.** i. Given a piece of combustion tubing, a U tube, a wash bottle and lime water, devise an experiment by which coal gas may be passed over heated red lead to show that water and carbon dioxide are formed. Burn the excess of coal gas issuing from the tube.

***Exp.** ii. Heat a mixture of copper oxide and charcoal and prove that the oxide is reduced to copper and that carbon dioxide is formed.

***Exp.** iii. Make a small hollow in a piece of charcoal and place in it either (*a*) litharge, (*b*) black oxide of copper, or (*c*) silver nitrate. Heat the substance strongly with the blowpipe flame. The metal is recovered chiefly by the reducing action of carbon monoxide formed over the red-hot charcoal.

106. Reduction of Nitric Acid. We have already shown in § 104 that nitric acid and nitre are oxidizing agents. Nitre yields oxygen on heating. Nitric acid if heated in a test tube boils away. But if the acid is decomposed by dropping it on a red-hot plate oxygen is set free.

To liberate oxygen from nitric acid.

Exp. i. Heat a little sawdust on a sand tray and let a few drops of nitric acid drop on the hot tray. The sawdust burns in the oxygen liberated.

Exp. ii. Having clamped a churchwarden pipe in a sloping position so that the mouthpiece dips into a pneumatic trough, heat the stem red-hot, and slowly drop strong nitric acid into the bowl of the pipe (Fig. 80). Collect the gas liberated and prove that it is oxygen.

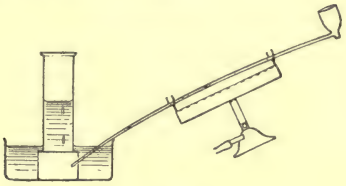


Fig. 80.

To liberate hydrogen from nitric acid.

Exp. Pour a very weak solution of nitric acid (3 parts acid to 100 parts water) on magnesium turnings in a small flask fitted with a delivery tube and collect the gas liberated in a test tube by water displacement: prove that it is hydrogen. This experiment requires considerable care to obtain the right conditions[1].

107. Nitric oxide and Nitrogen peroxide. It is evident then that nitric acid contains *oxygen*, which it readily liberates, and also *hydrogen*: the one an oxidizing and the other a reducing agent. As a rule when an acid dissolves a metal, the metal displaces the hydrogen of the acid, but, we recollect (§ 54) that reddish coloured fumes appeared, and *not* hydrogen, when nitric acid was poured on zinc and copper. How is this to be explained?

We must suppose that hydrogen is about to be liberated, but in the presence of the strong oxidizing agent nitric acid the hydrogen is oxidized to form water. It is a case of 'nascent hydrogen' and 'nascent oxygen' reacting together; the nascent hydrogen is displaced by the metal while the removal of nascent oxygen from some of the nitric acid causes the liberation of new gases, chiefly nitric oxide and nitrogen peroxide.

[1] A 10 per cent. solution of nitric acid acting on magnesium yields an explosive mixture of oxides of nitrogen and hydrogen.

108. To prepare nitric oxide.

Exp. i. In the flask that you used for preparing hydrogen, place about 15 grams of *copper*, with about 20 c.c. of water: connect to the pneumatic trough. Pour gradually strong nitric acid through the funnel on to the copper (Fig. 81). Reddish fumes rise which soon become colourless, and the solution turns blue as the copper dissolves. Collect three cylinders of the gas.

Fig. 81.

Exp. ii. Show that a lighted taper is extinguished by the gas.

Exp. iii. Show that phosphorus **[danger]**, heated strongly, burns in the gas. (Use a deflagrating spoon.) Nitrogen gas is liberated, and phosphoric anhydride (an acid forming oxide, § 51) is formed. Test with litmus solution.

Exp. iv. Into the third cylinder standing in the pneumatic trough *slowly* bubble oxygen. A red heavy gas is formed which dissolves in water. The water rises slowly in the jar. This red gas is **nitrogen peroxide**: its solution is acid.

To prove that nitric oxide contains nitrogen and oxygen pass a slow current of the gas through a long combustion tube containing copper and heated red-hot (cf. Fig. 62). Notice that the copper turns *black*. Collect the gas issuing from the end of the tube by water displacement. It puts out a taper, does not burn, does not turn lime water milky and is inert to phosphorus. We conclude that it must be *nitrogen*, but it is difficult to prove this without finding its density. If dry hydrogen is passed over the blackened copper, *water* is formed and the copper is restored. This proves that the nitric oxide converted copper to copper oxide (black) and therefore must have contained oxygen.

The blue liquid yields crystals of copper nitrate. When effervescence ceases in the nitric oxide generating flask, filter the blue solution from copper and evaporate to a small bulk and set it to cool in the desiccator. Beautiful blue crystals of copper nitrate form. Collect these crystals and dry them.

109. Effect of heat on a metallic nitrate. We have found that nitre (potassium nitrate) yields oxygen on being

heated. The nitrates of the heavy metals such as lead and copper yield the oxide of the metal, peroxide of nitrogen and oxygen.

Exp. Carefully warm the blue crystals in a dish to expel water of crystallization; then heat the anhydrous copper nitrate strongly in a test tube and prove that peroxide of nitrogen and oxygen are set free and that copper oxide is left.

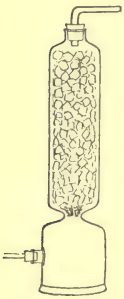


Fig. 82.

The nitrates of silver and of mercury yield the metal because the metallic oxide is decomposed by heat.

110. Oxidation of Ammonia. If a current of ammonia gas (§ 57), dried by passing it through a 'tower' filled with quicklime (Fig. 82), is led over black copper oxide, heated to redness, *water* and *nitrogen* are formed. The water may be condensed in a cooled U tube and the nitrogen may be collected by water displacement.

Questions on Chapter XVI.

1. Explain what is meant by 'oxidizing' and 'reducing' agents and give instances of each.

2. What happens when sulphur dioxide is passed into Condy's fluid? Give reasons.

3. Chlorine will not act as a bleaching agent unless moisture is present. How do you account for this?

4. Some parts of a gas flame exert an oxidizing action and other parts a reducing action on substances placed in them. Explain as clearly as you can what this means and give examples in illustration. (O. J.)

5. How can it be proved that nitric acid contains (*a*) hydrogen, (*b*) oxygen?

6. In what way do you account for the fact that when copper is acted upon by nitric acid hydrogen is not evolved?

7. Describe, with sketch of apparatus, how nitric oxide may be prepared. Why do the fumes in the collecting jar first appear red, then colourless?

8. Three gas jars contain respectively nitric oxide, oxygen and a mixture of equal parts of oxygen and nitrogen.

By what tests would you identify each?

9. Describe in as many ways as you can how lead oxide and copper oxide may be prepared.

10. How would you prove by experiment that ammonia gas contains nitrogen and hydrogen? Draw a diagram of the apparatus you would use.

CHAPTER XVII.

PETROLEUM, COAL GAS AND FLAME.

111. Petroleum (Rock oil) is found in the more recently formed and also in the ancient rocks which have been laid down in the sea by wastage from the land (sedimentary rocks). It occurs chiefly in South Russia, Upper Burmah, Pennsylvania, Ohio, California and some districts round the Gulf of Mexico. It is probably the product of decay of marine animal and vegetable remains under the pressure of the strata lying above.

It is obtained by boring a hole or sinking a well, sometimes to great depths, into the oil-bearing strata. Occasionally the oil wells up to the surface under great pressure accompanied by the release of *marsh gas.*

Petroleum has not a constant boiling point. It is separated by *fractional distillation* (§ 29) into petroleum ether (b.p. under 70° C.), petrol (b.p. over 70° but under 90°), benzene, paraffin oil (b.p. over 150°), and paraffin wax. All these substances are compounds of carbon and hydrogen.

Petrol and oil engines depend for their motive power on the expansion caused by the heating of gases produced by exploding the vapour of the petrol or oil with air. The petrol or oil is introduced into the cylinder of the engine, where the explosion takes place, in the form of vapour, the liquid being converted into gas by means of the *carburetter.* The mode of ignition is usually by the electric spark.

112. Coal Gas. Coal is another product of ancient vegetable life. In England it is found lying between the 'new red sandstone' (above) and the millstone grit and mountain limestone (below). *Anthracite* coal contains about 95 °/₀ carbon. Wigan, Newcastle and *Cannel* coal yield a large volume of inflammable gases on being heated out of contact with air. Students should visit gas works, where they will see coal heated strongly in fire-clay or iron *retorts*. The gases driven off are cooled and purified. Coal tar and impurities such as ammonia and sulphur compounds are removed and the gas is stored in large holders (*gasometers*). The process may be shown in the laboratory as follows:

Exp. Coal dust is heated in a hard glass tube and the products of 'dry distillation' are passed through the apparatus shown in Fig. 83. The names of the substances removed from the gas are given below in the diagram.

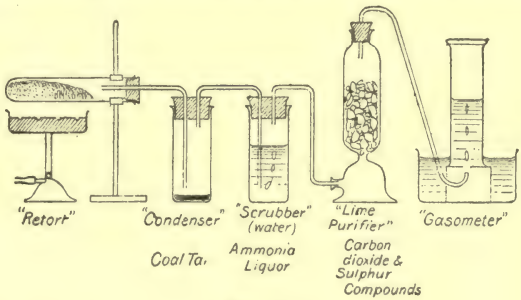


Fig. 83.

Coal tar is a thick alkaline liquid and contains many organic substances, including *aniline*, the basis of the valuable dyes, discovered by an Englishman, but manufactured chiefly in Germany.

Coal gas contains approximately 45 °/₀ hydrogen, 35 °/₀ marsh gas, 15 °/₀ carbon monoxide and 5 °/₀ ethylene. *Marsh gas* and *ethylene* are compounds of carbon and hydrogen; the latter compound gives coal gas its value as an illuminant.

Producer gas and Water gas are made by admitting air and steam into retorts where cheap coal is heated. A mixture of gases suitable for heating and manufacturing processes is obtained. A greater percentage of carbon monoxide, with nitrogen as an impurity, is present.

Wood gas may be obtained by distilling *sawdust* in an apparatus somewhat similar to that shown in Fig. 83, but without the 'scrubber' and 'lime purifier.'

Wood tar is acid in reaction and contains creosote, acetone, acetic acid and wood spirit.

Fire damp, the dangerous gas occurring in coal-mines, consists chiefly of marsh gas (methane). Mixed with air, 'fire damp' is explosive and forms '*after damp*'; the latter is dangerous because it contains the highly poisonous gas, carbon monoxide.

Explosions in coal-mines, as well as in flour and cotton mills, are often due to coal dust, or other finely divided organic matter, suspended in the air, igniting (by oxidation) sufficiently rapidly to produce a sudden burst of flame.

113. Flame is the region where combination is taking place between two or more gases, the change being accompanied by heat and light. In § 95 we burned hydrogen and coal gas in chlorine. We are familiar with the flames of coal gas, candles and lamps. Heat must be first applied before the vapours or gases will combine with the surrounding atmosphere. The lowest temperature at which flame ensues is called the **temperature of ignition.**

Exp. Show that a glowing splinter will ignite vapour of carbon bisulphide but will not ignite coal gas.

114. Mutual combination in flame.

Exp. i. Allow a small but steady stream of oxygen (or air) to issue from a glass jet held vertically and connected with an aspirator containing the gas into which water is running from the tap. Over this jet bring a jar containing

burning hydrogen. The oxygen is seen burning in the hydrogen. (Fig. 84.)

Exp. ii. If our atmosphere were coal gas, we could produce a flame by leading in air or oxygen through the gas pipes. Let us show this in

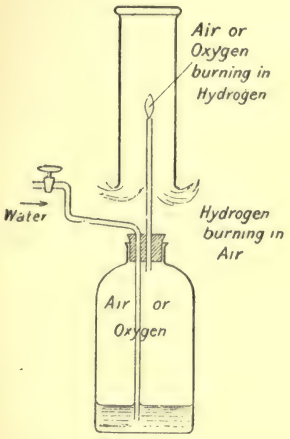


Fig. 84.

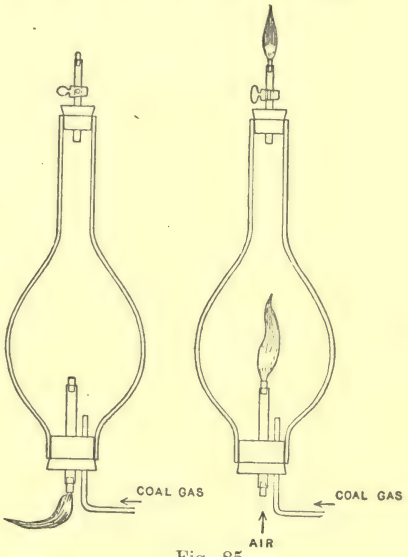


Fig. 85.

miniature. In Fig. 85 we have a lamp glass filled with coal gas. The tap at the top is first closed and the gas is lit at the bottom as it issues into the air. The tap is next opened and a light applied, when the lower flame turns inside out, runs back up the tube and *the air burns in an atmosphere of coal gas.*

115. The Oxy-hydrogen Flame. The simplest flame is shown with the ordinary 'blow-through' jet, used in a lime-light lantern. A flame of hydrogen is lit in the air and a current of oxygen is blown through. A small cone of unburnt oxygen is seen in the centre of the flame; round this cone is a zone of flame where combination is taking place between hydrogen and oxygen from inside and outside the zone.

The Bunsen Flame.

***Exp.** Open the holes at the bottom of the tube, turn on the gas very slightly and apply a lighted match to the top of your 'Bunsen.' Most probably there will be a slight explosion and the flame will '*strike back.*' The gas will burn inside at the bottom of the tube. The tube was filled with an explosive mixture of air and coal gas and the wave of ignition travelled down the tube more rapidly than the mixed gases ascended. Re-light your Bunsen and identify the parts of the flame by aid of Fig. 86. Insert a glass tube, held aslant, into the cone of unburnt gas and light the gas you 'tap off.'

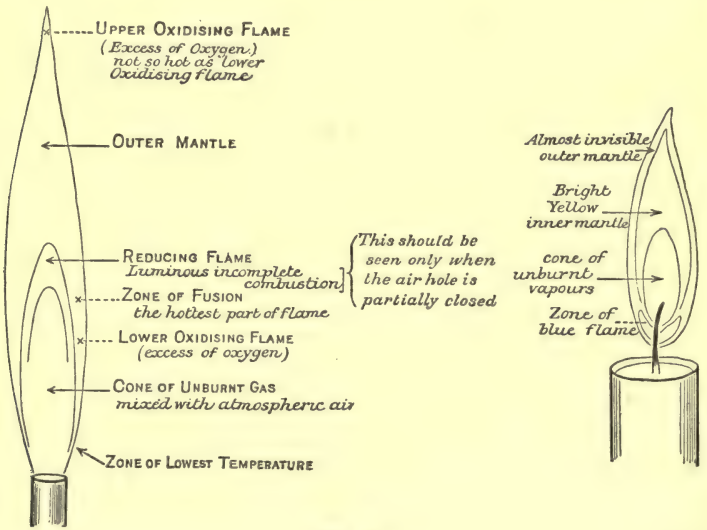

Upper Oxidising Flame
(Excess of Oxygen)
not so hot as Lower Oxidising Flame
Outer Mantle
Reducing Flame
Luminous incomplete combustion
Zone of Fusion
the hottest part of Flame
Lower Oxidising Flame
(excess of oxygen)
Cone of Unburnt Gas
mixed with atmospheric air
Zone of Lowest Temperature
This should be seen only when the air hole is partially closed
Almost invisible outer mantle
Bright Yellow inner mantle
cone of unburnt vapours
Zone of blue flame

Fig. 86.

The Candle Flame.

Exp. Fit a glass jet, by means of a cork, to a small test tube containing a few grams of paraffin wax. Heat the wax and light the 'candle gas' as it issues from the tube.

When you light a candle, the first process is the conversion of wax to vapour; this vapour is then ignited. Light a candle and identify the parts

of the flame shown in Fig. 86. Try to 'tap off' and burn some of the 'candle gas' from the inmost cone.

When you 'blow out' the candle, you cool the gases below their temperature of ignition. The flame may be extinguished by surrounding the wick with a small coil of copper wire which cools the flame by conducting the heat rapidly away (Fig. 87, A).

This experiment may be modified with the Bunsen flame.

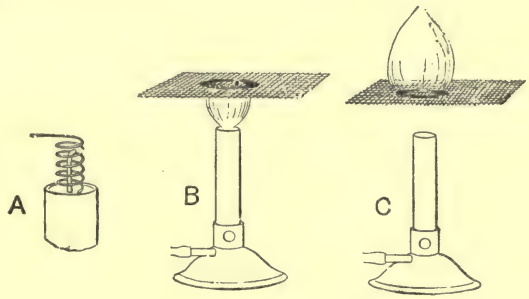


Fig. 87.

***Exp.** (i) Bring a piece of wire gauze of fine mesh on to the top of the Bunsen flame (Fig. 87, B). The heat is conducted away by the gauze and unburnt gases, cooled below their **ignition point,** issue through it. These may be lit above the gauze.

(ii) Turn off the gas and place the wire gauze on the top of the tube. Turn on the gas and light it above the gauze (Fig. 87, C). Raise the gauze and note that there is no flame immediately above the tube. Explain this.

Davy's Safety Lamp. In the simplest form of miner's safety lamp, a flame burns inside a small cage made of wire gauze. In the later forms, glass replaces the lower part of this cage, so that more light is transmitted. If the lamp is upset, the air supply is cut off and the lamp goes out. The substitution of glass also reduces the risk of the flame being blown through the lamp.

Exp. Lower a lighted safety lamp into a large beaker containing 3 or 4 c.c. of methylated ether, which rapidly evaporates and makes an explosive

mixture with air. The ether vapour is seen burning *inside* the lamp but the flame cannot penetrate the wire gauze. Often the lamp is extinguished by the carbon dioxide and steam formed by the burning ether.

116. Luminosity of Flame. Both a Bunsen and an oxy-hydrogen flame may be rendered luminous:

(*a*) by introducing a **solid** which may be heated to *incandescence, e.g.* a piece of platinum wire; a mantle of thorium oxide (Welsbach light); solid particles of carbon sprinkled over the flame or introduced with the air supply; a piece of quicklime heated white hot (lime-light):

(*b*) by introducing **heavy hydro-carbon vapour** such as benzene or petrol.

Non-luminosity of the Bunsen Flame may be produced:

(*a*) by **cooling** (the ordinary flame gives less light if a cold rod of copper is suddenly placed in the flame):

(*b*) by **diluting** the gas with an inert gas such as nitrogen or carbon dioxide:

(*c*) by **oxidation,** combustion being more complete and more rapid as the amount of oxygen present is increased:

(*d*) by **reducing** the **pressure** of the atmosphere.

Questions on Chapter XVII.

1. What is 'petrol'? How is it employed for driving engines?

2. What elements are contained in coal? What becomes of them when coal is burnt in an open fire and when it is distilled in a retort as in gas works? (O. J.)

3. What are the gases contained in coal gas? Sketch an apparatus to collect any gases which may be formed when a jet of coal gas burns in air and describe how their nature may be proved. (O. J.)

4. How could you prove that air will burn in coal gas?

5. What are 'water gas,' 'fire-damp,' 'after-damp'?

6. Sketch and describe the Davy safety lamp, and state its advantages and disadvantages.

7. What is the difference between an ordinary gas flame and that of a Bunsen burner? If both burnt the same amount of gas would they give out the same amount of heat, and if not, why not? (O. J.)

8. How may the luminosity of a flame be increased or diminished?

9. What is meant by 'temperature of ignition'? Explain what happens when (*a*) a *cold* coil of wire, (*b*) a *hot* coil of wire, is put over the wick of a lighted candle.

CHAPTER XVIII.

GENERAL CHARACTER OF METALS. EQUIVALENTS.

117. Physical Properties. The more obvious physical properties have been given in § 50. *Metals* are opaque, malleable, and ductile elements, each possessing a characteristic colour and lustre. They are heavier than water (except sodium and potassium), they are all solid (except mercury), and they conduct heat and electricity.

Chemical Properties. In §§ 56 and 60 we found that the oxides of metals are *basic*, *i.e.* they neutralize acids to form salts. Solutions of metallic oxides are *alkaline*. Metals, which dissolve in an acid, displace the hydrogen of the acid and form a salt. We shall find also in this chapter, that in electrolysis of salts, metals are deposited at the *negative electrode* or kathode.

Native Metals. Metals which do not oxidize readily occur in the free state (native) in nature and consequently have been known to man from the earliest times. Copper, tin, iron, gold, silver, lead and mercury were used by the ancients. Sodium, potassium and calcium have been discovered comparatively recently, by aid of the electric current. Aluminium, during the past few years, has been isolated on a large scale by electrolytic processes.

118. Recovery of metals from their Ores.

***Exp.** Repeat Exp. iii, § 105, and obtain lead from galena and copper from azurite.

The Blast Furnace. The ores of several of the metals, the commonest of which is iron ore, are reduced to metal in a

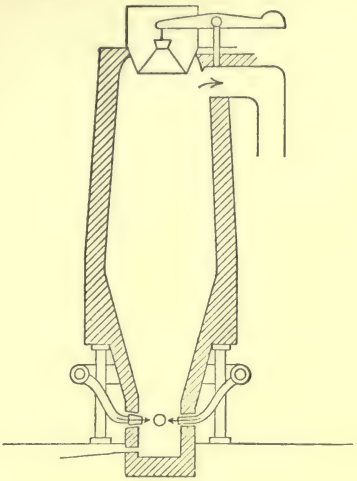


Fig. 88.

furnace of fire brick built in cylindrical form (Fig. 88). The ore (let us take for example *clay ironstone*) is previously *calcined*, *i.e.* roasted to expel water and carbonic acid gas. The oxide of iron thus obtained contains much earthy matter, silica, an acid oxide, which will combine at a high temperature with a basic oxide such as lime. The calcined iron ore, mixed with coke and limestone, is put in at the top of the furnace and a blast of hot air is pumped in at the bottom. The coke burns, forming carbon monoxide, which reduces the iron ore to *iron*. The limestone is decomposed into lime and carbon dioxide; the lime combines with the silica forming *slag* and the carbon dioxide is reduced to carbon monoxide on coming in contact with red-hot coke. The slag trickles down the furnace, carrying the reduced iron with it, and collects above the molten iron, thus protecting the latter from being oxidized by the blast of hot air which enters at the bottom

of the furnace. The iron is 'tapped' from the hearth and is run into moulds in a bed of sand. This variety is called 'pig iron.' The waste furnace gases, containing much carbon monoxide, are burned and utilized for heating the blast.

Electric Furnaces. A much higher temperature (2000°—3000° C.) may be obtained by passing a strong electric current through the ores. The great heat is partly due (*a*) to the electric arc and partly (*b*) to the resistance to the passage of electricity. In addition, decomposition of the compounds melted in the furnace may be caused by electrolysis (see §§ 119 and 120).

An electric furnace consists essentially of a box or crucible lined with plumbago (graphite), a conductor of electricity (*NNN*

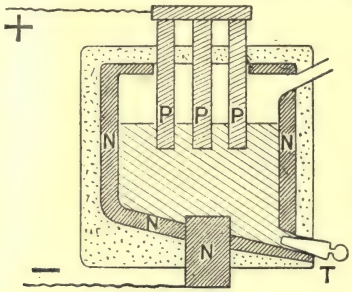


Fig. 89.

in Fig. 89), which is connected with the negative terminal of a powerful dynamo. The positive terminal is joined to carbon rods *PPP* which dip into the mineral matter contained in the crucible. The molten material may be 'tapped' by means of a plug *T*.

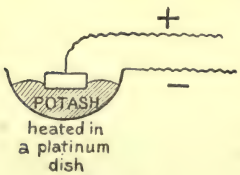


Fig. 90.

119. Electrolysis. Refer to § 67, where we obtained hydrogen and oxygen by passing a current of electricity through a dilute solution of sulphuric acid.

Sir Humphry Davy, in 1807, using a platinum dish as negative electrode

(kathode) passed electricity through fused potash and soda, and obtained for the first time the metals potassium and sodium (Fig. 90).

120. Electrolysis of Salts in solution.

Exp. i. **Copper Sulphate.** Pass an electric current through a solution of copper sulphate, using two pieces of platinum as electrodes. Notice that *copper* deposits on the negative electrode (kathode) and that *oxygen* is liberated at the positive electrode (anode); the solution meanwhile becomes less coloured. Next substitute a piece of copper for the platinum anode. The solution remains coloured, copper is still deposited on the kathode and the copper anode dissolves.

Exp. ii. **(1) Dilute Sulphuric acid, (2) Copper Sulphate, (3) Silver Nitrate, (4) Hydrochloric acid, (5) Sodium Chloride, in the same circuit.** Connect five electrolytic cells containing the above solutions 'in series' with

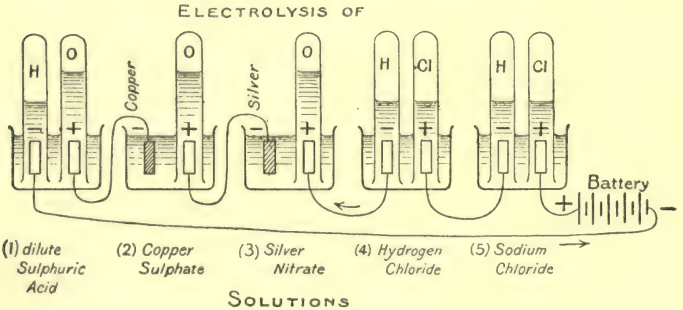


Fig. 91.

a battery of about six accumulators (Fig. 91). Hydrogen in equal volumes is liberated at the kathode in cells (1), (4) and (5). After several hours have elapsed and the solutions are thoroughly saturated, equal volumes of chlorine will collect at the anode in cells (4) and (5) and half this volume of oxygen at the anodes of cells (1), (2) and (3). Notice that in cells (1) and (4) where acids are electrolysed hydrogen is liberated at the kathode. In cells (2) and (3) the *metals* copper and silver are deposited on the kathode; it is reasonable to suppose that the *metal* sodium would be deposited at the kathode in cell (5) but immediately it is set free it reacts with water to form hydrogen and sodium hydrate. The latter may be detected by its action on litmus paper.

121. Equivalents. By careful experiment we can weigh the quantity of metal or gas liberated by the same current of electricity. If the kathodes in cells (2) and (3) are weighed before and after passing the current, the increase in weight gives the weight of copper and silver deposited. We know the relative densities of hydrogen, oxygen and chlorine and we know the weight of 1 c.c. of hydrogen = ·00009 gram; therefore by measuring the volumes (S.T.P.) of the gases liberated we can calculate their weights. The results are as follows:

If the weight of hydrogen set free = 1,
then " " oxygen " = 8,
" " chlorine " = 35·5,
" " copper " = 31·5,
" " silver " = 108.

Now these are the proportions by weight in which these elements combine together; thus,

In 9 parts (by weight) of water there are 1 part of hydrogen, 8 parts of oxygen (§ 65).

In 36·5 parts (by weight) of hydrogen chloride there are 1 part of hydrogen, 35·5 parts of chlorine.

In 39·5 parts (by weight) of black copper oxide there are 31·5 parts of copper, 8 parts of oxygen.

These weights are said to be **Equivalent.**

Definition of Chemical Equivalent. The chemical equivalent of an element is the **weight of the element which combines with or displaces 1 part by weight of hydrogen** (or 8 parts by weight of oxygen, &c.).

Revision Exercises.

Ex. 1. We already know how to collect and measure the volume of hydrogen displaced when a known weight of metal dissolves in acid; devise an experiment to show that the equivalent of magnesium = 12. Use the aspirator and remember that 1000 c.c. of hydrogen (S.T.P.) weigh ·09 gram.

Ex. 2. Confirm this result by showing that magnesium combines with oxygen in the proportions of 12 : 8.

Ex. 3. Reduce a known weight of dry oxide of copper (black) in a stream of hydrogen or coal gas and show that 31·5 parts by weight of copper combine with 8 parts of oxygen.

Ex. 4. Confirm this result by carefully dissolving a gram of copper in a small quantity of nitric acid in a weighed dish; evaporate slowly and finally heat strongly until the nitrate is converted to oxide of copper.

Ex. 5. Devise a method of proving the equivalent of zinc to be 32·5.

122. Displacement of a metal in solution by another metal. We have already noticed that, when iron filings are added to copper sulphate solution, copper is deposited and iron dissolves (§§ 20 and 98). This reaction goes on quantitatively, the metals replacing each other in the proportions of their equivalent weights.

***Exp.** i. **To find the equivalent of iron** assuming the *equivalent of copper* to be 31·5. Weigh about 1 gram of *dry* 'reduced' iron into a beaker; pour on to the iron a hot solution of copper sulphate made by dissolving about 7 grams of crystals in 100 c.c. of water. Stir well for five minutes and filter through two balanced filter papers (§ 25). Wash all the copper on to the filter with hot water and finally rinse the filter paper with rectified spirit. Dry the filter over the cone; separate the papers and weigh the copper. Calculate the weight of iron which would displace 31·5 grams of copper. [Eq. Wt. Iron = 28.]

***Exp.** ii. **Find the equivalent of silver** given the equivalent of magnesium = 12. Weigh accurately 0·1 gram magnesium and leave it to stand in a solution of *silver nitrate* (2 grams of crystals in 50 c.c. of water). Proceed as above, Exp. i. [Eq. Wt. Silver = 108.]

Questions on Chapter XVIII.

1. Describe the process of recovering iron from iron ore by means of the blast furnace.

2. Give a sketch of an electric furnace and explain its action.

3. Define 'electrolysis.' What is the effect of passing an electric current through copper sulphate solution using (*a*) platinum electrodes, (*b*) copper electrodes?

4. How were sodium and potassium first isolated?

5. What substances are produced during the electrolysis of solutions of (*a*) silver nitrate, (*b*) sodium chloride? Give reasons.

6. Define the term 'chemical equivalent.' Sketch and describe the apparatus you would employ for finding the chemical equivalent of zinc.

7. If excess of sulphuric acid is added to 13 grams of zinc what weight of hydrogen will be evolved?

CHAPTER XIX[1].

DALTON'S ATOMIC THEORY. SYMBOLS, FORMULAE, EQUATIONS. CHEMICAL CALCULATIONS.

123. The Atomic Theory. To explain the *Law of Fixed Proportions* (§ 84) and the *Law of Multiple Proportions* (§ 86), John Dalton, of Manchester, in 1804, formulated the following *theory*. He revived the ancient Greek idea that matter is not capable of being divided indefinitely. Imagine, for instance, that a piece of silver is cut into two pieces, and each piece again into two and so on a million times, it is supposed that finally a small particle is obtained which cannot be cut or divided, *i.e.* an **atom** (from a Greek adjective meaning *indivisible*).

Matter is supposed to be made up of these atoms.

All atoms of the same element have equal weights.

The weight of an atom of an element is called its *atomic weight.*

Chemical combination takes place between atoms.

The smallest particle of an element or compound capable of free existence is called a **molecule.** If elements in the free state existed as single atoms and compounds were formed by the combination of single atoms then the atomic weight and the equivalent weights of elements would be the same. But there are reasons to suppose that *molecules* of elements generally contain more than one atom. Thus the smallest particles of oxygen in the air are called molecules of oxygen; each is supposed to consist of two atoms.

[1] For fuller treatment of the subjects of this chapter, reference should be made to a larger text-book, *e.g. Inorganic Chemistry*, by E. I. Lewis, Cambridge University Press.

In 1808, **Gay-Lussac,** a Frenchman, discovered that **gases combine in simple proportions by volume.** We know that

1 vol. Hyd.	+	1 vol. Chlorine	form	2 vols. Hydrogen Chloride

(§ 93).

If we assume, as was suggested by **Avogadro** (1811), that **equal volumes of gases** under the same conditions of temperature and pressure **contain equal numbers of molecules,** it follows that

n molecules of hydrogen + n molecules of chlorine form $2n$ molecules of hydrogen chloride.

Therefore, dividing throughout by n,

1 molecule of hydrogen + 1 molecule of chlorine form 2 molecules of hydrogen chloride.

Each molecule of hydrogen and of chlorine helps to form two molecules of hydrogen chloride. To explain this further it is simplest to assume that each of these molecules contains 2 *atoms.*

Again

2 vols. hydrogen	+	1 vol. oxygen	form	2 vols. steam

(§ 68),

hence, two molecules of hydrogen, each supposed to contain two atoms, combine with one molecule of oxygen, containing two atoms, to form two molecules of water, each containing two atoms of hydrogen and one atom of oxygen. This may be represented as follows:

Let an atom of hydrogen = (H), and an atom of oxygen = [O],

then $\begin{matrix}(H)(H)\\(H)(H)\end{matrix} + [O][O] = \begin{matrix}(H)[O](H)\\(H)[O](H)\end{matrix}$,

2 molecules of hydrogen + 1 molecule of oxygen = 2 molecules of water.

Hence we obtain the definition of an **atom**: *the smallest particle of an element which can enter into or be expelled from a compound.*

If the atomic weight of hydrogen = 1,
then its molecular weight = 2.

We have found that 1 part by weight of hydrogen combines with 8 parts by weight of oxygen (§ 65); if the above equation represents the mode of combination of hydrogen and oxygen, then 4 atoms of hydrogen weighing 4 combine with 2 atoms of oxygen weighing $4 \times 8 = 32$, and, therefore, each atom of oxygen must weigh 16.

The molecule of water, on this supposition, weighs 18, being made up of two atoms of hydrogen (2×1) combined with one atom of oxygen (16).

Valency. The equivalent weight of chlorine is 35·5. Its atom weighs 35·5 times the weight of an atom of hydrogen (= 1). Its atomic weight and equivalent weight being the same, an atom of chlorine has a combining value or *valency* of a *single* equivalent weight, *i.e.* the atomic weight of chlorine = 1 × its equivalent weight. An atom of chlorine can combine with one atom of hydrogen and is said to have a *valency of one* and to be *monovalent* (Greek *monos* = one).

The equivalent weight of oxygen is 8. Its atomic weight = 16. The atomic weight of oxygen = 2 × its equivalent weight, and therefore the atom of oxygen weighing 16 can combine with $\frac{16}{8} = 2$ atoms of hydrogen; hence its *valency* is said to be **2.** Oxygen is *di*valent.

The equivalent weight of nitrogen is 4·67. Its atomic weight = 14. The atomic weight of nitrogen = 3 × its equivalent weight, and therefore the atom of nitrogen weighing 14 can combine with $\frac{14}{4 \cdot 67} = 3$ atoms of hydrogen; hence the *valency* of nitrogen = **3.** Nitrogen is *tri*valent.

The **Valency** of an atom is the *number* of hydrogen (or univalent) atoms which it can combine with (or displace).

The atomic weight of an element = its valency × its equivalent weight.

Dulong and Petit in 1819 discovered that the *atoms of all elements have the same capacity for heat*; or, **atomic weight × specific heat = a constant = 6 (approx.).**

$$\therefore \text{ the approximate atomic weight} = \frac{6}{\text{specific heat}}.$$

This value is sufficiently near to give the valency and hence we can find the exact atomic weight from the equivalent weight which may be determined with accuracy.

Example. Aluminium. Specific Heat = ·21.

$$\therefore \textit{Approximate} \text{ Atomic Weight} = \frac{6}{\cdot 21} = 28{\cdot}5,$$

but Equivalent Weight = 9 (found *with accuracy* by displacement of Hydrogen).

$$\therefore \text{ Valency of Aluminium} = \frac{28{\cdot}5}{9} = 3 \ (\textit{approx.}).$$

$$\therefore \textit{Exact Atomic Weight} = \text{Equivalent Weight} \times \text{Valency}$$
$$= 9 \times 3 = \mathbf{27}.$$

124. Symbols. Each element is represented by a *symbol*, usually a letter, which is not only used as a short method of writing the name of the element, but also represents the *weight* of the *atom* of the element. Thus:

H =	an atom of	hydrogen	weighing	1,
O =	,,	oxygen	,,	16,
Cl =	,,	chlorine	,,	35·5,
Ca =	,,	calcium	,,	40,
C =	,,	carbon	,,	12.

125. Formulae. A molecule of an element or compound containing several atoms united together may be represented by formulae, thus:

(1) A molecule of water = H_2O and consists of 2 atoms of hydrogen and 1 of oxygen; its total molecular weight = 18[1].

(2) A molecule of calcium chloride = $CaCl_2$ and consists of 1 atom of calcium weighing 40 and 2 atoms of chlorine weighing 35·5, its total molecular weight therefore = $40 + (35{\cdot}5 \times 2) = 111$.

(3) Two molecules of calcium carbonate = $2CaCO_3$
$= 2[40 + 12 + (16 \times 3)] = 2 \times 100 = 200$.

126. Equations.

(1) The formation of water by combination of hydrogen and oxygen (§ 123) may be represented as follows:

$$2H_2 + O_2 = 2H_2O,$$
$$2(1 \times 2) + 16 \times 2 = 2[(1 \times 2) + 16],$$
$$4 + 32 = 36 \text{ parts by weight.}$$

(2) The decomposition of calcium carbonate, similarly:

$$CaCO_3 = CaO + CO_2,$$
$$40 + 12 + (16 \times 3) = (40 + 16) + (12 + \overline{16 \times 2}),$$
$$100 = 56 + 44 \text{ parts by weight}$$

(3) The action of hydrochloric acid on marble, similarly:

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2,$$
$$100 + 73 = 111 + 18 + 44 \text{ parts by weight.}$$

127. Calculations. Weight. If we require to find the weights of two or more substances reacting together it is necessary to know (*a*) their chemical formulae, (*b*) the atomic weights of the elements composing the substances, (*c*) the chemical equation representing the reaction in respect of weight.

Question i. What weight of quicklime (CaO) is obtained by heating 25 tons of marble ($CaCO_3$)?

Equation. $CaCO_3 = CaO + CO_2$,
$100 = 56 + 44$ (by weight);

[1] N.B. The number of atoms are represented by figures below the line: the weights of the elements are to be *added* together: a number placed in front of a formula multiplies the molecular weight by that number.

$\therefore$ 100 tons marble yield 56 tons quicklime,

$\therefore$ 1 ton marble yields $\frac{56}{100}$ ton quicklime,

$\therefore$ 25 tons marble yield $\frac{56 \times 25}{100} = 14$ tons quicklime.

Question ii. What weight of caustic soda (sodium hydrate, NaOH) is required to neutralize 15 tons of hydrochloric acid (HCl) in solution and what weight of sodium chloride (NaCl) will be formed? Atomic Weight of Sodium = Na = 23.

Equation. $NaOH + HCl = NaCl + H_2O$,

$(23 + 16 + 1) + (1 + 35{\cdot}5) = (23 + 35{\cdot}5) + (\overline{1 \times 2} + 16)$,

$40 + 36{\cdot}5 = 58{\cdot}5 + 18$ (by weight);

$\therefore$ 36·5 tons of hydrochloric acid neutralize 40 tons of caustic soda and produce 58·5 tons of sodium chloride.

$\therefore$ 15 tons of hydrochloric acid neutralize $\frac{40 \times 15}{36{\cdot}5}$ tons of caustic soda and produce $\frac{58{\cdot}5 \times 15}{36{\cdot}5}$ tons of sodium chloride.

Volume. The molecular weight of a gas expressed in grams occupies 22·2 litres (S.T.P.).

Thus 2 grams of hydrogen (H_2) occupy 22·2 litres S.T.P.

44 grams of carbon dioxide (CO_2) occupy 22·2 litres S.T.P.

Question i. What volume of hydrogen S.T.P. will be obtained by dissolving 7 grams of zinc (Zn = 65) in sulphuric acid ($= H_2SO_4$)?

$Zn + H_2SO_4 = ZnSO_4 + H_2$,

65 grams of zinc displace 2 grams or 22·2 litres of hydrogen;

$\therefore$ 1 gram of zinc displaces $\frac{22{\cdot}2}{65}$ litres of hydrogen S.T.P.,

$\therefore$ 7 grams of zinc displace $\frac{22{\cdot}2 \times 7}{65} = 2{\cdot}39$ litres of hydrogen S.T.P.

128. The following Equations represent some of the reactions of the experiments mentioned in the corresponding §§.

23. $2Na + Cl_2 = 2NaCl$.

34. $2Hg + O_2 \rightleftarrows 2HgO$.

35. (i) $2Mg + O_2 = 2MgO$. (ii) $2Cu + O_2 = 2CuO$.

37. $6PbO + O_2 \rightleftarrows 2Pb_3O_4$.

38. $2BaO + O_2 \rightleftarrows 2BaO_2$.

46. $2KClO_3 = 2KCl + 3O_2$.

48. (i) $P_4 + 5O_2 = P_4O_{10}$. (ii) $S + O_2 = SO_2$. (iii) $C + O_2 = CO_2$.

54. $KNO_3 + H_2SO_4 = KHSO_4 + HNO_3$.

57. $2NH_4Cl + CaO = CaCl_2 + H_2O + 2NH_3$.

58. $NaOH + HCl = NaCl + H_2O$.

59. (ii) $KOH + HNO_3 = KNO_3 + H_2O$.
(iii) $NH_4 . OH + HCl = NH_4 . Cl + H_2O$.
$NH_4Cl \rightleftarrows NH_3 + HCl$.

60. (i) $CuO + H_2SO_4 = CuSO_4 + H_2O$.
(ii) $PbO + 2HNO_3 = Pb(NO_3)_2 + H_2O$.

62. $Zn + H_2SO_4 = ZnSO_4 + H_2$. 64. $PbO + H_2 = Pb + H_2O$.

66. (i) $2K + 2H_2O = 2KOH + H_2$.
(iii) $Ca + 2H_2O = Ca(OH)_2 + H_2$.

72. $CaCO_3 \rightleftarrows CaO + CO_2$.

75. $CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$.

76. $CaCO_3 + H_2O + CO_2 \rightleftarrows CaH_2(CO_3)_2$.
(i) $CaH_2(CO_3)_2 + Ca(OH)_2 = 2CaCO_3 + 2H_2O$.

77. (i) $CaCl_2 + Na_2CO_3 = 2NaCl + CaCO_3$.

85. $CO_2 + C = 2CO$.

88. (*a*) $C + 2CuO = 2Cu + CO_2$; (*b*) $H_2 + CuO = Cu + H_2O$.

91. $NaCl + H_2SO_4 = NaHSO_4 + HCl$.

93. $ZnO + 2HCl = ZnCl_2 + H_2O$.

94. $MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2$.

97. $Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2$.

99. (ii) $SO_2 + H_2O + O = H_2SO_4$. (iii) $2SO_2 + O_2 = 2SO_3$.

100. $BaCl_2 + H_2SO_4 = BaSO_4 + 2HCl$.

101. $FeS + 2HCl = FeCl_2 + H_2S$.

104. $H_2O + Cl_2 + H_2SO_3 = 2HCl + H_2SO_4$.

108. (i) $3Cu + 8HNO_3 = 3Cu(NO_3)_2 + 4H_2O + 2NO$.

109. $2Cu(NO_3)_2 = 2CuO + 4NO_2 + O_2$.

110. $4NH_3 + 3O_2 = 6H_2O + 2N_2$.

122. (i) $CuSO_4 + Fe = FeSO_4 + Cu$.
(ii) $Mg + 2AgNO_3 = Mg(NO_3)_2 + 2Ag$.

129. Table of Symbols, Atomic Weights, Densities, &c., of Elements.

	Symbol	Atomic Weight (A)	Equivalent Weight (E)	Valency = $\frac{A}{E}$	Density (water = 1)	Coeff. Linear Expansion ·000...	Specific Heat
Non-Metals							
Carbon	C	12	3	4	2—3·5		·15—·24
Chlorine	Cl	35·5	35·5	1	35·5 (Density Hyd. = 1)		
Hydrogen	H	1	1	1	1 (Density Hyd. = 1)		
Nitrogen	N	14	4·67 or 2·8	3 or 5	14 (Density Hyd. = 1)		
Oxygen	O	16	8	2	16 (Density Hyd. = 1)		
Phosphorus	P	31	10·33 or 6·2	3 or 5	1·8—2·3		·2
Sulphur	S	32	8 or 5·33	4 or 6	2		·18
Metals							
Aluminium	Al	27	9	3	2·6	023	·21
Calcium	Ca	40	20	2	1·4		·145
Copper (*Cuprum*)	Cu	63·5	31·7	2	8·95	0168	·094
Iron (*Ferrum*)	Fe	56	28	2	7·7	011	·113
Lead (*Plumbum*)	Pb	207	103·5	2	11·3	029	·031
Magnesium	Mg	24	12	2	1·7	027	·246
Mercury (*Hydrargyrum*)	Hg	200	100	2	13·6		·033
Potassium (*Kalium*)	K	39	39	1	·86		·166
Silver (*Argentum*)	Ag	108	108	1	10·5	0192	·056
Sodium (*Natrium*)	Na	23	23	1	·97		·293
Zinc	Zn	65	32·5	2	7·14	0292	·094

Questions on Chapter XIX.

1. What do you mean by 'atom,' 'molecule,' 'molecular weight,' 'chemical symbol'?

2. Write a short account of Dalton's Atomic Theory and show how it explains the Laws of Fixed and Multiple Proportions.

3. If 16 grams of carbon are burnt in oxygen what weight of carbon dioxide will be produced?

4. What weight of calcium carbonate is required to neutralize 18·25 gm. of hydrochloric acid in solution and what weight of calcium chloride will be formed?

5. What volume will be occupied by 7 grams of carbon monoxide at S.T.P.?

6. What volume of sulphur dioxide will be produced if 4 grams of sulphur are burnt in oxygen?

7. If 2 grams of water are electrolyzed what volume of hydrogen will be obtained?

8. What weight of chalk will be produced if 556 c.c. of carbon dioxide are passed into lime water?

REVISION QUESTIONS SELECTED FROM THE OXFORD AND CAMBRIDGE EXPERIMENTAL SCIENCE PAPERS.

Paper A.

1. Describe the chemical and physical differences between a piece of charcoal and a lump of lead.

2. How would you show the presence in the air of a room of (*a*) water vapour, (*b*) carbon dioxide, (*c*) solid particles?

3. Briefly describe three methods of preparing hydrogen from water. What volume of hydrogen could be obtained from 10 grams of water (1 litre of hydrogen weighs 0·09 gm.)?

4. What experiments both qualitative and quantitative would you make to ascertain the changes which occur when a piece of chalk is strongly heated?

5. What are the relations between salt, chlorine, and hydrochloric acid? Describe how you would prove that hydrochloric acid contains hydrogen.

6. Describe the appearance of roll sulphur. What other forms of sulphur are known? Describe experiments to show that sulphur dioxide, sulphuretted hydrogen, and sulphuric acid all contain sulphur.

7. Some metals are often found as metals in nature, *e.g.* gold, copper, silver; others, such as zinc, are never found except in combination. Can you explain why this is the case? Mention some other elements which you would not expect to find free and give your reasons for your expectation.

8. Describe how you could most easily obtain a sample of pure water (*a*) from lime water, (*b*) from a solution of caustic soda, (*c*) from dilute hydrochloric acid.

Paper B.

1. Describe and compare the physical and chemical properties of any three common solid elements and one compound of each.

2. Describe three simple experiments by which it can be shown that more than one substance is present in the atmosphere. Name the four chief substances which are present and say what you know of their amounts.

3. How is the composition of water ascertained?

4. Describe experiments to show that pure chalk and calcium carbonate are chemically identical.

5. By what reactions may hydrogen and chlorine be obtained from ordinary hydrochloric acid? Describe the properties and uses of this acid.

6. Describe the behaviour of sulphur when heated and the properties of the substance formed when it burns in air. For what purpose is sulphur used?

7. In what forms are the metals potassium, tin, aluminium, and copper most commonly met with in nature? Describe how you would obtain any *one* of the metals mentioned from a compound containing it.

8. What takes place when (*a*) phosphorus, (*b*) sodium, (*c*) iron is exposed to the air? Will the substances formed dissolve in water? If so, what are the properties of the solutions thereby obtained?

Paper C.

1. Describe experiments by which you could show the chief chemical and physical properties of three of the following substances: Copper, silver, sulphuric acid, soap, soda, sugar.

2. Describe a method of preparing oxygen. How would you prove that oxygen is a constituent of (i) air, (ii) chalk?

3. State your reasons for believing that water is a chemical compound and not a mixture of hydrogen and oxygen. What happens when water is added to (*a*) sodium metal, (*b*) strong sulphuric acid, (*c*) quicklime?

4. Describe experiments by which you could show that (*a*) carbon dioxide is not an element, (*b*) it is more soluble in water under high pressure than under low pressure.

5. How is chlorine prepared? What are its properties?

6. Describe some of the properties of either nitric acid or hydrochloric acid or sulphuric acid, showing how to distinguish the acid you select from the other acids.

7. When either zinc or chalk is added to a solution of hydrochloric acid a gas is given off. Compare the properties of the gases obtained in the two cases.

Have we any grounds from other experiments for deciding whether the gas is derived from the solid or the liquid?

8. What changes occur when the following substances are strongly heated in air: (*a*) blue crystals of copper sulphate, (*b*) gypsum, (*c*) chalk, (*d*) sugar, (*e*) lead, (*f*) manganese dioxide?

State in each case whether the substance gains or loses in weight, and in the case of (*a*), (*b*), (*c*) the effect of moistening the residue with water.

Paper D.

1. Give the names of eight different metals, and roughly classify them according to the ease with which they (*a*) undergo oxidation in air, (*b*) dissolve in dilute acids.

2. Give three examples showing that when substances combine chemically they do so in definite proportions by weight, and three other examples shewing that in the case of gases they combine also in definite and simple proportions by volume.

3. Describe several experiments by which the properties of hydrogen can be illustrated, stating in each case the inferences to be drawn from the experiments you describe.

4. Explain how to find by experiment what volume of carbon dioxide is formed on burning a given weight of pure charcoal in air.

5. Describe the methods by which chlorine and hydrochloric acid gas are made from common salt. Describe two experiments which show the chief properties of chlorine and of hydrochloric acid.

6. Write out the names of the elements contained in sulphuric acid. What is the effect of adding the acid to (*a*) chalk, (*b*) common salt, (*c*) copper metal? How would you carry out each experiment?

7. What are the metals gold, iron, zinc, lead most commonly used for and why?

8. Describe and explain if you can what happens when (*a*) limestone is strongly heated, (*b*) water poured on lime, (*c*) hydrochloric acid treated with an oxidizing agent, (*d*) turpentine burns.

Paper E.

1. By what properties could you recognize the following substances: (*a*) water, (*b*) sulphur, (*c*) sand, (*d*) oxygen?

2. How would you prepare and collect carbon dioxide if you were given chalk and hydrochloric acid? Give a sketch of the apparatus you would use.

3. If 50 c.c. of hydrogen and 55 c.c. of oxygen are exploded together in a eudiometer what weight of water will be formed and what weight of uncombined gas left? A litre of hydrogen weighs 0·09 gm. If the temperature of the eudiometer were kept at 100° C. what would be the total volume of gas after the explosion?

4. What happens if we pass a stream of hydrochloric acid gas through lime water? If the amount of gas absorbed by the solution is known it is possible to calculate in what quantities certain other substances are formed. Why is this and what important principle is thereby illustrated?

5. What two gases are commonly used for bleaching purposes? Explain in each case the chemical nature of the changes occurring and state why the presence of water is necessary. How do watery solutions of these gases affect one another and how could you show that some chemical change takes place after they are mixed?

6. What are the chief bodies present in coal gas? What are the substances formed when it is burnt in the ordinary gas flame? How would you demonstrate their presence?

7. A solid substance when heated gives off a gas. By what experiments would you endeavour to identify the gas?

ANSWERS.

Ch. IV.

4. (*a*) Sand 15·68 gm. Nitre 84·32 gm.
   (*b*) Sand 90·9 gm. Chromate 9·1 gm.
6. 36 gm. 25°·5 C.
7. (*a*) 29·54 gm. (*b*) 36·24 gm.
8. 21·4 gm. 9. 19·42 gm.

Ch. VI.

6. 66·6 %. 7. 27 gm.

Ch. VIII.

3. 25·51 gm. 4. 4·864 gm.

Ch. X.

8. *A* 15 %. *B* 36·84 %.

Ch. XII.

3. 56·08 % quicklime. 43·92 % carbon dioxide.
6. 9·9 gm.

Ch. XIX.

3. 58⅔ gm.
4. 25 gm. carb. 27·75 gm. chlor.
5. 5560 c.c. 6. 2780 c.c.
7. 2471⅓ c.c. 8. 2·5 gm.

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